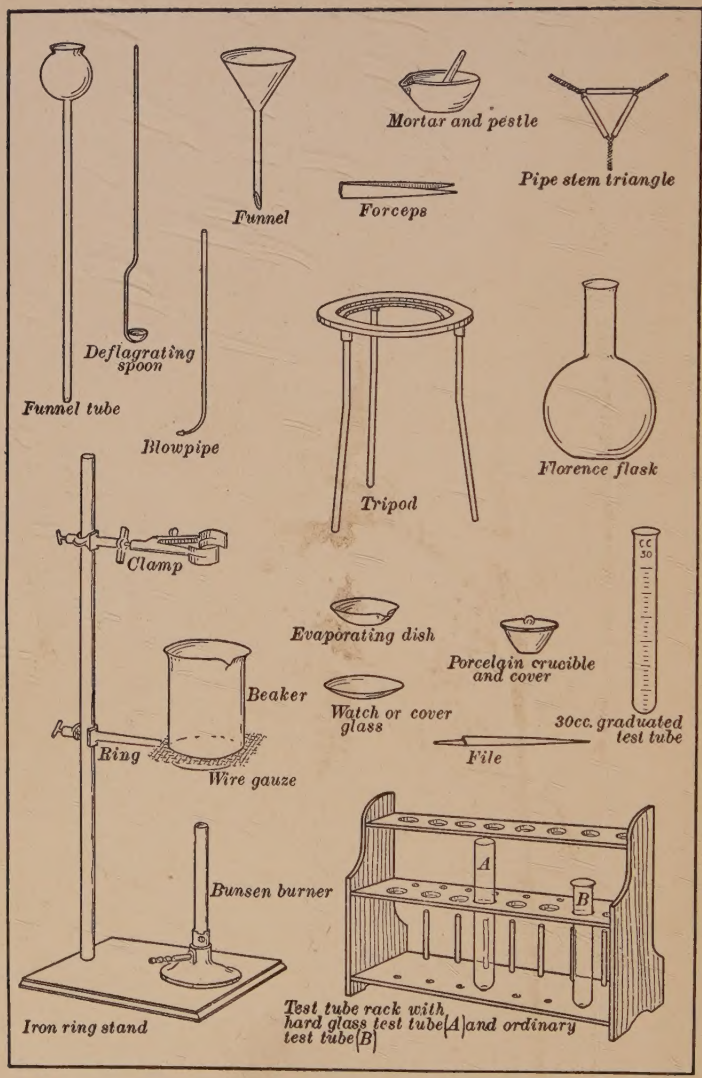
LABORATORY PRACTICE IN CHEMISTRY

McPHERSON AND HENDERSON

James M Toy
Nov. 24, 1929.
Senior Class.
Laboratory



Some apparatus included in the student's outfit

LABORATORY PRACTICE IN CHEMISTRY

EXERCISES TO ACCOMPANY

"CHEMISTRY AND ITS USES"

BY

WILLIAM McPHERSON

AND

WILLIAM EDWARDS HENDERSON

GINN AND COMPANY

BOSTON · NEW YORK · CHICAGO · LONDON

ATLANTA · DALLAS · COLUMBUS · SAN FRANCISCO

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528.7

The Athenæum Press
GINN AND COMPANY · PROPRIETORS · BOSTON · U.S.A.

PREFACE

In selecting the experiments to accompany the text "Chemistry and its Uses" the authors have made every effort to limit their choice to such as have been subjected to extended and successful trial in a large number of secondary schools in which ordinary conditions prevail. The authors have relied not only upon their own experience with large numbers of students in elementary chemistry but also upon the coöperation and criticism of a number of teachers in secondary schools throughout the country. It is a pleasure to acknowledge in this place the assistance so willingly rendered us.

It is clearly impossible, however, to make a selection that will meet with the approval of all teachers. Indeed, an author can hardly hope to secure the entire approval of any one teacher. Each one works under certain limitations of time, equipment, and local interests, and each has his individual training and viewpoint. An experiment regarded very highly by one is often considered rather fruitless by another. One desires to stress the applications of chemistry and another prefers to adhere to more fundamental principles. And all this is exactly as it should be with enthusiastic teachers.

In order that each teacher may have some choice, more experiments are included in the book of laboratory exercises than can be performed by all the class within the time ordinarily allotted to laboratory work. Some of the exercises are marked "Optional," while parts of others are printed in fine type, indicating that they may also be omitted if time is not available.

FOREWORD TO THE STUDENT

The laboratory is a place for seeing, but it is equally a place for thinking. The laboratory course will include a certain amount of work that is largely mechanical, such as bending and sealing glass tubes, but in the main it is headwork rather than handwork that counts. Each experiment is designed to bring out an important fact or to illustrate some principle. Be sure you grasp the meaning of the experiment before you start upon it. Study the directions with the same care that you would give to an assignment in the textbook for recitation.

Keep your desk neat and clean. A glance at a student's work desk will usually tell an experienced teacher exactly what kind of work is being done at it. A dirty desk, littered with unnecessary articles and with apparatus awkwardly put together, almost surely indicates a lack of understanding of the work in hand and a willingness to accept without question any thing the book says. A clean desk and well-connected apparatus indicate clear thinking and real joy in the work. Be determined to prove the truth of things for yourself.

Get ready for an experiment before you start it. In addition to your burner, have at hand all the materials called for at the beginning of the directions for each experiment. There will be four or five bottles of the most-used reagents on your desk, and the others will be on the shelf for general use.

Be cautious about reagents. Some of the substances you will handle are more or less poisonous. Some are inflammable, while others are explosive under certain conditions. Many attack the skin and occasion painful wounds. There are many

reagents in the laboratory that cannot be mixed with safety. Take time to read labels correctly; you cannot afford to make mistakes. There is absolutely no danger if you follow directions, but you must exercise the same care as does a mechanic in a machine shop or a driver of an automobile.

Be economical. Reagents are expensive, and someone must pay the bills. Experiments nearly always go better with small quantities than with large ones, and are more quickly carried out. Every good chemist works with the smallest quantities possible—partly to save money, but more largely to save time and learn more. Stick to the quantities specified in the notes. Actually weigh or measure these quantities until you can guess very closely how much a gram or a cubic centimeter really is.

Make your notes correspond to the questions. The questions to which written answers are to be given are indicated by letters of the alphabet in parentheses; thus, (*a*). Always place the same letter before your reply to connect it with the proper question. Note the abbreviations frequently used in the directions: grams (g.), centimeters (cm.), cubic centimeters (cc.), liters (l.). The abbreviation R. S. means that a given reagent is to be sought on the general *reagent shelf* and not upon your own desk. Some exercises are marked "Optional," while portions of others are printed in fine type to indicate that these portions are also optional. The best students of the class will find time to perform at least some of these optional experiments. The list of apparatus and materials required for optional experiments is included in the list at the beginning of the exercise, but is inclosed in parentheses.

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LABORATORY PRACTICE IN CHEMISTRY

EXERCISE 1

THE METRIC SYSTEM (SEE APPENDIX). METHODS OF MEASURING THE LENGTH, VOLUME, AND WEIGHT OF VARIOUS OBJECTS

Apparatus. Graduated test tube; 2 balances, one for weighing from 0.5 g. to 1000 g. and the other from 0.01 g. to 50 g. or 75 g.; watch glass; 100-cc. beaker.

Materials. 5 or 6 g. common salt.

NOTE. Exercise 1 is inserted for the benefit of any students who may not have studied physics or who for any reason may not be familiar with the metric system or with the process of weighing small objects.

FIG. 1. Ten-centimeter scale

1. Length. By means of the 10-centimeter scale (Fig. 1) measure and (*a*) record the length of various pieces of apparatus included in your outfit, as a test tube, a file, and a blowpipe. (*b*) What is the diameter of your filter paper? Estimate the lengths of various objects, as a pencil, a test tube; then measure. Continue until you can approximate the lengths of such objects.

FIG. 2. Graduated test tube

2. Volume. By means of a graduated test tube or cylinder (Fig. 2) measure in cubic centimeters and (*c*) record the volumes of various test tubes, beakers, and flasks included in your outfit. In reading off the amount of the liquid in a graduated tube, always read from the lower part of the *meniscus*; that is, the curved surface of the liquid (Fig. 3).

3. Weight. The platform balance (Fig. 4) is adapted for weighing objects that weigh from 0.5 g. to 1000 g. It is not so

sensitive as the chemical balance (Fig. 5) or the hornpan balance (Fig. 6), either of which is sensitive to 0.01 g. These smaller balances, however, while very sensitive, can only be used in weighing objects up to 50 g. or 75 g. at the most. The arms of the chemical balance (Fig. 5) can be raised or lowered by turning the screw *D*. When the balance is not in use the arms should always be lowered so that the pans rest on the top of the base of the balance.

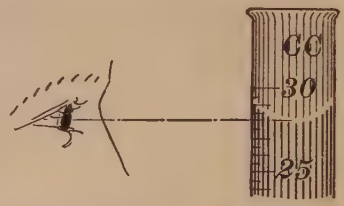

Fig. 3. Taking the reading of the volume of a liquid in a graduated test tube

Before using a balance the two arms must be in exact equilibrium. In the case of the platform balance, if the arms are in equilibrium, then the pointer *A* (Fig. 4) will point directly to the zero mark on the scale *B* when the pans are at rest; if it does not, bring it to this position by adjusting the screw *C*. In the more sensitive balance (Fig. 5) the pointer *A* must swing equal distances on either side of the zero point marked on the scale *B* when the pans move up and down; if it does not, make it do so by turning the screw *C*.

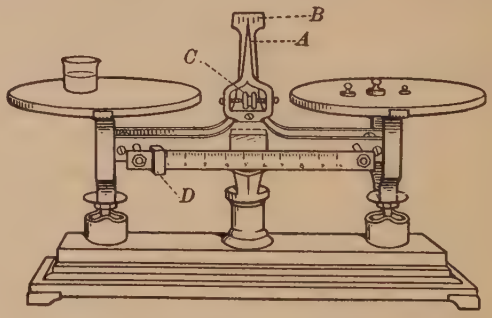

Fig. 4. Platform balance for weighing objects up to 1000 g.

When the balance is adjusted, the object to be weighed is placed on one pan (the left is more convenient) and weights (Figs. 7 and 8) are added to the other until the arms are in equilibrium. In the case of the platform balance the equivalent of 10 g. or any fraction of this weight is obtained by simply moving the

slide *D* (Fig. 4) along the scale. When the arms are in equilibrium the weights are carefully noted and recorded. After a little practice one can approximate closely the weight of the object and thus know about what weights must be added. *Care must be taken to read and record the weights accurately.* It is a good plan to check the readings of the weights by first noting the vacant places in the box in which the weights are kept.

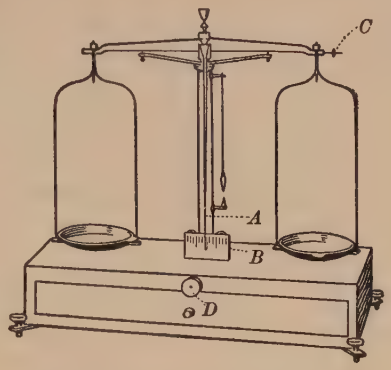

Fig. 5. Chemical balance for accurately weighing small objects up to 50g. or 75g.

It is not advisable here to give detailed directions in regard to the use of the balance; such information will be furnished by the instructor. It must be remembered, however, that a balance is a delicate instrument and must be given the same care as a watch. The objects weighed must be clean and dry, and the weights must be handled only with small forceps (Fig. 9). Under no condition must chemicals be placed directly on the scalepan. Such substances as sugar and salt may be weighed out on paper; others, which act upon paper, must be weighed in a small beaker or on a watch glass.

Fig. 6. Hornpan balance. This can be used in place of the chemical balance, but it is not so convenient, although less expensive

After studying the construction of the balance and the method of using it, weigh and (*d*) record the weights of various small objects, as a porcelain crucible, a watch glass.

(*e*) What is the approximate weight of a nickel five-cent piece?

Accurately balance a watch glass on the scalepan (Fig. 5), using either the weights or, better, a small pill box and fine sand, and weigh out on this exactly 5.2 g. of common salt.

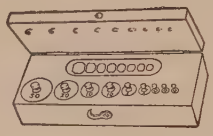

FIG. 7. Weights used with the chemical balance

FIG. 8. Weights used with the platform balance

Clean and dry a small beaker and balance it on the scalepan; then remove it from the pan and pour into it, as nearly as possible, 10 cc. of distilled water (measured in the graduated tube included in your laboratory outfit). Reweigh, and (*f*) note the weight of the water. (*g*) How do you account for the fact that it does not weigh exactly 10 g.?

FIG. 9. The proper way of handling weights

EXERCISE 2

THE BUNSEN BURNER

Apparatus. A Bunsen burner and rubber tube; wing-top burner.
Materials. Wooden splints (1 cm. × 12 cm. is a convenient size); white cardboard (10 cm. square).

1. Construction. The Bunsen burner is a form of apparatus used for burning gas and is the common burner employed in the laboratory. It consists of the tube *A* (Fig. 10), screwed into the base *C*. The tube has two small openings near its lower part. A small band *B*, provided with similar openings, fits around the lower part of the tube in such a way that the openings of the tube may be closed or kept open by turning the band. Gas is admitted through *D* by means of rubber tubing connected with the gas pipe.

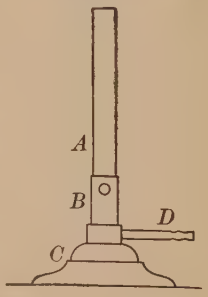

FIG. 10. A Bunsen burner

Unscrew the tube and examine the different parts of the burner (Fig. 11); then put them together again. Turn on the gas and at once ignite it by holding a lighted match 4 or 5 cm. above the top of the tube. The supply of gas should be adjusted so as to give a flame about 10 cm. high. The gas flowing through the tube mixes with the air drawn in through the openings in the lower part of the tube and burns with an almost nonluminous flame. If the band is adjusted so as to close the openings, the flame becomes luminous. Always use the nonluminous flame unless otherwise directed.

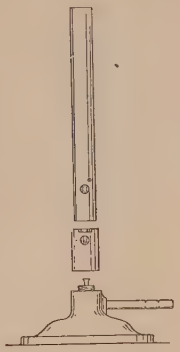

FIG. 11. The parts of a Bunsen burner

Turn off the gas, slip over the end of the tube the so-called "wing-top" burner (*A*, Fig. 12), and again ignite the gas. Notice that the flame is spread out. This form of the flame is always used in bending glass tubes, as is described in Exercise 3.

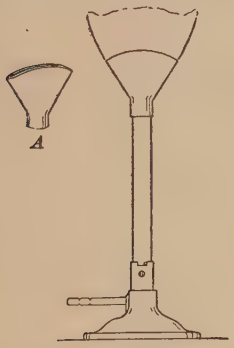

FIG. 12. Wing-top burner (*A*), and also a Bunsen burner fitted with a wing-top burner

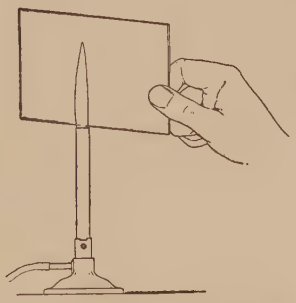

FIG. 13. Testing the relative temperatures of the different parts of a flame by holding a cardboard in the flame

2. Determining the relative temperatures of the flame. Hold a small wooden splint horizontally in the base of the Bunsen

flame for two or three seconds and (*a*) note the results. In the same way determine the relative temperatures of various parts of the flame. Turn the gas down until the flame is 7 or 8 cm. in height; then quickly thrust a piece of white cardboard *vertically* through the center of the flame, the lower end of the cardboard resting against the top of the burner (Fig. 13). Remove the cardboard before it is ignited, and from the scorched portions note the relative temperatures of different parts of the flame. (*b*) Draw a diagram to illustrate your results.

EXERCISE 3

THE MANIPULATION OF GLASS TUBING

Apparatus. Burner and wing-top attachment; hard-glass test tube with cork to fit; cork-borers; triangular file; round file.

Materials. Glass tubing (soft), external diameter 6 mm.; about 30 cm. of glass rod.

1. To fit a tube with stopper and glass tube as shown in Fig. 14. In all operations requiring the application of a strong heat to glass, the heat must be applied gently at first; otherwise a strain is produced which is apt to break the glass. For the same reason highly heated glass must be cooled slowly.

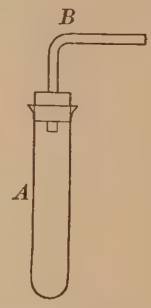

Fig. 14. A test tube fitted with cork and glass tube

From one of the lengths of soft-glass tubing cut a piece about 15 cm. in length. To do this, place the tubing on the desk and, holding it firmly with your hand, draw the edge of a triangular file across the point at which you wish to cut the glass. After the glass has been scratched, take the tube in the hands with the thumbs placed near together just back of the scratch (Fig. 15) and gently *pull* the glass apart, at the same time exerting a slight pressure with the thumbs. If the tube does not yield readily to a gentle pressure, a deeper scratch must be made. In the case of large

tubing it may be found necessary to file a groove around the glass. The edges of the cut tube will be sharp and should be rounded by being rotated in the tip of the Bunsen flame (Fig. 16).

Fig. 15. After the glass has been scratched with a file, it is pulled apart as shown

To bend the glass tubing, first heat it, at the point where you wish to bend it, in the luminous Bunsen flame spread out by means of the wing-top burner (Fig. 17). Hold the tube lengthwise in the flame, gently rotating it so that all sides may be equally heated. Continue the heating until the glass bends easily, then remove it from the flame and quickly bend it to a right angle *B* (Fig. 14). Great care must be taken to heat the tube uniformly; otherwise the bore of the tube will be contracted (*A*, *B*, Fig. 18), forming a bend which is not only unsightly but is easily broken.

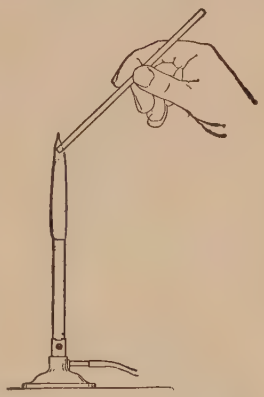

Fig. 16. Rounding the sharp ends of a glass tube or rod

Next select a good cork of such a size that the smaller end will just enter the hard-glass test tube. Soften the cork by rolling it between the desk and a block of wood. The next step is to insert the glass tube, prepared as directed above, into and through the cork. To do this, select a borer (*A*, Fig. 19) slightly smaller than the tube. Place the cork on a piece of cardboard on the desk and, holding the rod *B* firmly with your hand, cut through the cork with the borer, not by punching but by rotating the borer under gentle pressure (Fig. 19). Care must be taken to keep the borer at a right angle to the top and base of the cork. The hole should be straight and smooth.

The end of the glass tube is now moistened with water or vaseline and *gently* pushed through the cork by a screwlike motion. If the hole is too small to admit the tube when a gentle pressure is applied, it must be slightly enlarged with a round file. *It must be remembered that the glass is fragile and will break if more than a gentle pressure is applied; and if broken while grasped in the hand, a serious cut may result.* Now put the cork in the test tube and set the apparatus aside for use in preparing oxygen (Exercise 7).

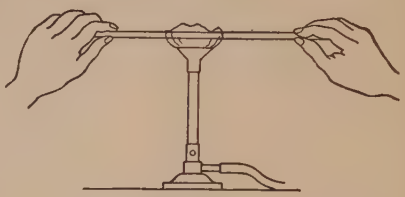

FIG. 17. Heating tube before bending it

FIG. 18. Poorly bent tubes

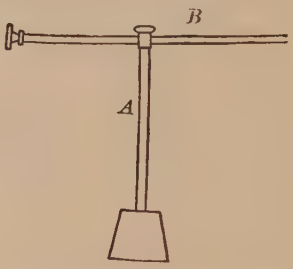

FIG. 19. Boring a hole through a cork stopper

2. To make glass stirring-rods. These should be about 15 cm. in length. Cut the glass rod into pieces of this length, using the method employed in cutting glass tubing. Round the ends of the rods by heating in the Bunsen flame (Fig. 16). Place the finished rods in the desk for future use.

EXERCISE 4

SOME FURTHER PRELIMINARY MANIPULATIONS

Apparatus. 2 beakers; graduated tube; glass stirring-rod; funnel; evaporating-dish; mortar and pestle.

Materials. Filter paper; about 1 g. each of sand and salt.

NOTE. Since a Bunsen burner and a ring stand are on each desk, they will not be included in the list of apparatus in subsequent exercises.

1. Pouring liquids from one vessel to another. In doing this care must be taken to prevent the liquid from running down the side of the vessel from which it is poured. A glass rod

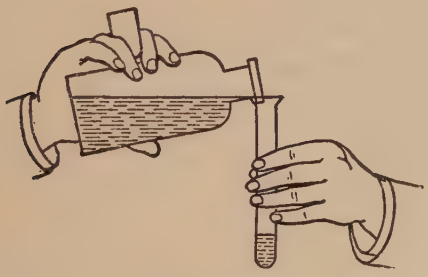

FIG. 20. Pouring liquid from a bottle into a tube

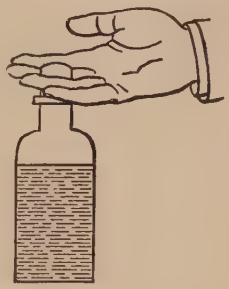

FIG. 21. Removing the stopper from a bottle

(prepared in Exercise 3) should be held lightly against the rim of the vessel, as shown in Fig. 23. The liquid will flow down the rod.

Fill a beaker with water and transfer it slowly to another vessel without using the glass rod; repeat, using the glass rod.

In pouring liquids from bottles a glass rod may be used, or the neck of the bottle may be placed lightly against the rim of the vessel into which the liquid is being poured (Fig. 20). This will prevent the liquid from running down the side of the

bottle. The stopper must never be laid on the desk. Catch it between the fingers, as shown in Fig. 21, leaving the hand free to grasp the bottle, as shown in Fig. 20.

2. Decantation; filtration. It is often necessary to separate a liquid from a finely divided solid which is suspended in it. This may be done by one of the two following methods:

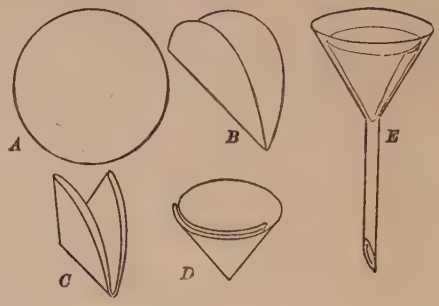

FIG. 22. The proper way of folding a filter paper and fitting it in a funnel

When the solid is heavy and readily settles to the bottom of the vessel, the liquid may be carefully poured off, or *decanted*. Sand suspended in water may be separated from the liquid in this way.

Usually, however, the solid will not readily settle or will do so only after long standing. In such cases the mixture is *filtered*, by pouring it on a filter paper, which allows the liquid to run through but retains the solid. The method of folding the filter paper and fitting it into a funnel is shown in Fig. 22. The paper (*A*, Fig. 22) is first folded along a diameter into halves (*B*); then, at right angles to the first fold, into quarters (*C*). The folded filter is then opened so as to form a cone (*D*), half of which is composed of three thicknesses of paper, and the remainder of one thickness. Now fit the cone into a funnel of such a size that the paper does not quite reach the top (*E*). The paper must fit the funnel

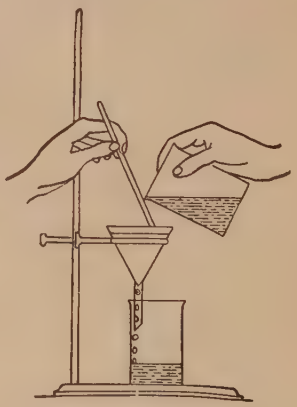

FIG. 23. Separating a solid from a liquid by filtration

accurately; if it does not, make it do so by varying the fold. Place the paper in the funnel and thoroughly wet it with water. After the water has run through, press the paper firmly against the sides of the funnel with the finger so as to remove any air bubbles between the paper and the glass. The filter is now ready for use (Fig. 23). The process of filtration not only enables us to separate liquids from solids but also certain solids from each other, as is illustrated in the following experiment:

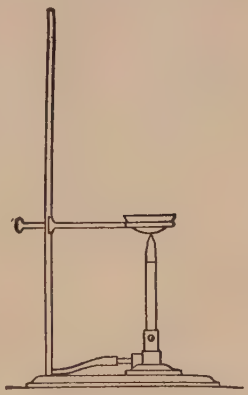

Fig. 24. Evaporating a liquid

3. Separating two solids from each other. Weigh out about 1 g. each of sand and common salt and mix them thoroughly by grinding them together in a mortar. Transfer the mixture to a small beaker, add 10 cc. of water, and stir the mixture for two or three minutes with a glass rod. Filter off the undissolved solid, collecting the *filtrate* (that is, the liquid that runs through the paper) in a beaker (Fig. 23). Pour the filtrate into an evaporating-dish. Support the dish on a ring stand (Fig. 24) and evaporate the liquid by a gentle heat. Withdraw the heat as soon as the water is evaporated. (*a*) What is the solid left on the filter paper? (*b*) What is the solid left in the evaporating-dish (taste)? (*c*) Can all mixtures of solids be separated by the above method?

EXERCISE 5

CHEMICAL COMPOUNDS; ELEMENTS; CHEMICAL ACTION (CHEMICAL CHANGE); CHEMICAL AFFINITY

Apparatus. Test tube (ordinary); hard-glass test tube.
Materials. Iron wire 10 cm. in length; wooden splints; 1 g. sugar; 1 g. mercuric oxide.

Caution. *In heating any substance in a tube care must always be taken never to point the tube toward yourself or anyone else; for in some cases the hot contents may be ejected from the tube and, if they come in contact with the body, may cause more or less serious injury. The correct way of heating a tube is shown in Fig. 25.*

FIG. 25. The correct way of heating a liquid in a tube

1. Hold a piece of iron wire in the Bunsen flame for a few seconds. (*a*) Is the iron changed? Examine the wire after it has cooled. (*b*) Have the properties of the iron been changed by the heating? (*c*) Has the iron undergone a chemical change?

Repeat the experiment, using a wooden splint. (*d*) Does the wood undergo a chemical change? (*e*) Contrast the action of heat on the iron and the wood.

2. Introduce 1 g. of sugar into a clean, dry test tube. This is done by placing the sugar near the end of a narrow strip of folded paper and introducing it into the tube held in a horizontal position, as shown in Fig. 26. If the paper is now turned over, the sugar is deposited in the bottom of the tube; the paper is then carefully withdrawn, leaving the sides of the tube perfectly clean.

Fold a piece of paper 12 or 15 cm. in length into the form of a band from 1 to 2 cm. in width, place the band about the upper end of the tube, and grasp the ends of the band between your thumb and forefinger, as shown in Fig. 25. The tube may now be heated without burning your fingers. Heat the sugar gently in the tip of the flame, as shown in Fig. 25, moving the tube about in the flame so as to heat the sugar uniformly. (*f*) Note carefully the changes that take place. Continue the heating as long as any changes occur and until a black solid is left. This solid is the element carbon. (*g*) Contrast the properties (taste, color, solubility in water) of the sugar and the carbon. (*h*) What is the source of the carbon? (*i*) What grounds have you for assuming that a chemical change has taken place?

FIG. 26. Introducing a small amount of solid into a tube

3. Introduce (Fig. 26) about 1 g. of mercuric oxide into a clean, dry, hard-glass test tube. Heat the oxide strongly, inserting a glowing splint from time to time down into the tube. Note that the splint bursts into a flame. This is caused by the invisible gas *oxygen*, which is evolved when the mercuric oxide is heated. Note the residue clinging in droplets on the sides of the tube near the bottom. This is the element *mercury* (*quicksilver*).

4. (*j*) Define the following terms, illustrating each by examples taken from the above experiments: *element, chemical compound, chemical change*, and *chemical affinity*.

EXERCISE 6

THE COLLECTION OF GASES; THE PREPARATION OF OXYGEN (PRELIMINARY)

Apparatus. Three 250-cc. wide-mouthed bottles; 3 pieces of window glass 10 cm. square; pneumatic trough; rubber tubing; 2 test tubes.

Materials. 4 g. potassium chlorate; 1 g. powdered manganese dioxide; wooden splints.

1. Collection of gases. Fill a 250-cc. wide-mouthed bottle with water. Cover its mouth with a glass plate, being careful to exclude all air bubbles. Hold the plate firmly in place, invert the bottle, and dip its mouth into the water in a pneumatic trough. Remove the glass plate. (*a*) Why does the water remain in the bottle? Now, without changing the position of the bottle, fill it with exhaled air by placing one end of a piece of glass or rubber tubing under the mouth of the bottle and blowing gently through the other end (Fig. 27).

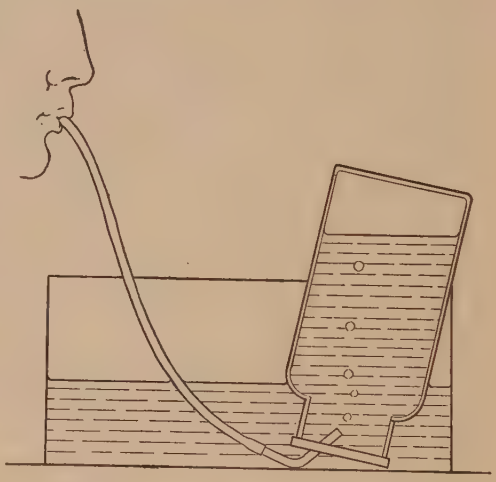

FIG. 27. Filling a bottle with exhaled air

Before the bottle, so filled, is removed from the trough, cover its mouth tightly with a glass plate. The bottle so covered may then be placed on the desk either right side up or in an inverted position, according to the weight of the gas.

Fill a bottle with exhaled air and then transfer the air to another bottle. (*b*) Draw a diagram to show a good method of doing this.

2. **Preparation of oxygen from mercuric oxide.** Recall 3, Exercise 5.

3. **Preparation of oxygen from potassium chlorate (preliminary experiment).** Select two test tubes of the same size and clean and dry them thoroughly. Into the one introduce 2 g. of potassium chlorate; into the other introduce 2 g. of potassium chlorate mixed intimately with 1 g. of manganese dioxide.

Now heat the contents of the two tubes gently, applying the flame so that both tubes are equally heated. Repeatedly thrust a glowing splint into each tube in order to detect any oxygen that may be evolved. (*c*) Note the results. (*d*) What effect has the manganese dioxide? (*e*) From which tube is the greater amount of oxygen evolved (consult text)?

EXERCISE 7

THE LABORATORY PREPARATION OF OXYGEN; PROPERTIES OF OXYGEN

NOTE. In some exercises certain experiments are optional. The paragraphs describing the optional experiments are in *fine print*. The apparatus and materials required for the optional experiments are given in the list, but the names are inclosed in parentheses (—).

Apparatus. Hard-glass test tube fitted with cork and delivery tube as shown in Fig. 28 (use apparatus constructed in Exercise 3, Fig. 14); four 250-cc. wide-mouthed bottles; 4 pieces of glass 10 cm. square; pneumatic trough; deflagrating-spoon; (evaporating-dish; funnel).

Materials. 6 g. potassium chlorate; 3 g. manganese dioxide; wooden splints; 1 g. of sulfur; picture-frame wire 20 cm. long; bit of cotton; (filter paper).

1. Usual laboratory method for preparing oxygen. Mix intimately on paper 6 g. of potassium chlorate and 3 g. of manganese dioxide. The presence of impurities in the materials

may lead to a serious explosion when heat is applied; hence *test a small portion of the mixture*, say 0.5 g., by heating it in a test tube. In the absence of impurities the oxygen is evolved quietly, unaccompanied by any very marked sparking in the materials.

If pure, transfer the remainder of the mixture to a hard-glass tube, insert the cork, and arrange the apparatus as shown

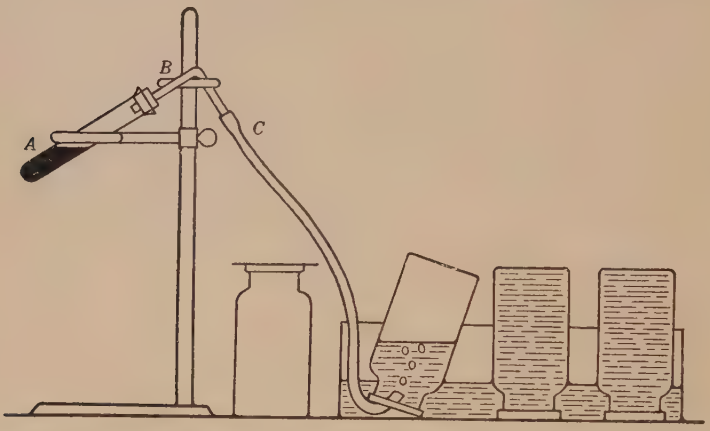

FIG. 28. Apparatus required for preparing oxygen by heating a mixture of potassium chlorate and manganese dioxide

in Fig. 28. The bottles must be completely filled with water and inverted in the water in the pneumatic trough. Now holding the burner in the hand, heat the mixture gently with a small flame, applying the heat at first to the upper part of the mixture. The flame must not strike the upper part of the test tube, as the cork may be burned. At first the heat expands the air, and a few bubbles of air escape; then the oxygen is evolved. Regulate the heat so as to secure a uniform and not too rapid evolution of the gas. Collect three or four 250-cc. wide-mouthed bottles of the gas. Before the heat is withdrawn, remove the cork from the tube. (*a*) What is the reason for this precau-

tion? Slip a piece of glass over the mouth of each bottle; then remove the bottle from the water and set it in an upright position on your desk, keeping the mouth of the bottle covered by the glass.

(*b*) What is the source of the oxygen? (*c*) What is the function of the manganese dioxide? Place the tube and contents aside for use in 3, below.

2. Properties and chemical conduct of oxygen. (*d*) Note the properties (color, odor, taste, solubility in water) of the gas. (The slight cloud that is often present when oxygen is prepared from potassium chlorate is due to an impurity and will disappear if the gas is allowed to stand over water.)

(*e*) Repeatedly thrust a glowing splint into a bottle of the gas and note the results.

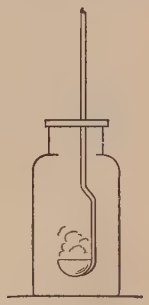

FIG. 29. Burning sulfur in oxygen

Heat some sulfur in a deflagrating-spoon until it begins to burn. (*f*) Note the color and size of the sulfur flame. (*g*) Now lower the burning sulfur into a bottle of oxygen (Fig. 29) and note the change.

Tip the piece of picture-frame wire with sulfur by wrapping a small bit of cotton about the end of the wire and dipping this into melted sulfur (for this purpose melt a little sulfur in a deflagrating-spoon). Ignite the sulfur by holding it in a Bunsen flame for an instant, and then thrust the end of the wire into a bottle of oxygen. (*h*) Record the results.

(*i*) Summarize the properties and chemical conduct of oxygen.

3. Separation of the compounds present in the residue left in the preparation of oxygen. Heat the tube containing the residue left in the preparation of oxygen until no more oxygen is evolved. (*j*) What two compounds are present in the residue (consult text)? One of these is soluble in water and the other insoluble. Devise a method for the separation of the two compounds and carry out the process. (*k*) Describe the method and record your results.

EXERCISE 8

A STUDY OF SOME OF THE CHANGES TAKING PLACE WHEN A SUBSTANCE BURNS

Apparatus. Porcelain crucible; pipestem triangle to support the crucible while being heated; small beaker; glass tube 12 or 15 cm. in length; 250-cc. wide-mouthed bottle; balance; (pneumatic trough; large wide-mouthed bottle or large beaker).

Materials. 2 or 3 g. finely powdered iron (use iron reduced by alcohol); candle; 10 cc. limewater (R.S.) (the abbreviation R. S. signifies that the reagent is found on the reagent shelf); (pellet of phosphorus size of a pea, to be obtained from instructor when needed; iron wire about 15 cm. long).

1. The burning of metals. Iron may be taken as a typical metal, since the metals in general burn in the same way as iron. Place 2 or 3 g. of powdered iron in a porcelain crucible and weigh to 0.01 g., (*a*) recording the weight. Now heat the crucible (Fig. 30) until the iron begins to glow (burn); then withdraw the heat. (*b*) Does the iron continue to burn? When the crucible is cool, reweigh. (*c*) Compare the weight of the product with that of the unburned iron and account for any change in weight.

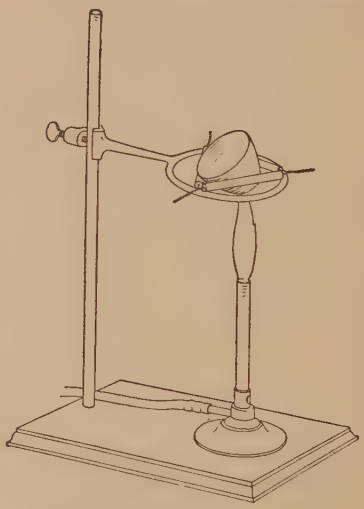

FIG. 30. Burning iron powder

2. The burning of a candle. Pour 5 cc. of clear limewater into a small beaker, and by means of a glass tube gently blow exhaled air through the liquid (Fig. 31). Note that the liquid soon becomes cloudy,

or "milky," in appearance. This change is due to the action of the colorless and odorless gas known as *carbon dioxide* (which is present in the exhaled air) and constitutes a good test for this gas.

Hold a cold, dry, wide-mouthed bottle over a candle flame as shown in Fig. 32. Note the film of moisture (water) collecting on the bottle. (*d*) What is the source of this water? After one or two minutes remove the bottle quickly and pour into it 5 cc. of clear limewater. Place the palm of the hand tightly over the mouth of the bottle and shake the contents. (*e*) Note any change in appearance of the limewater. (*f*) What conclusions do you draw from the experiment? (*g*) How do you account for the fact that a burning candle gradually disappears, while iron, on burning, increases in weight?

FIG. 31. Blowing exhaled air through limewater

3. The burning of phosphorus. (***Precaution.***) *Phosphorus must be kept and handled only under water; otherwise it may ignite and serious results follow.* Cover the bottom of a pneumatic trough with water to a depth of 2 or 3 cm. On the water float a porcelain crucible. Ask the instructor to place in this crucible a small piece of phosphorus the size of a pea.

Ignite the phosphorus by touching it with the end of a hot wire or file (Fig. 33), and quickly invert over the crucible a wide-mouthed bottle or beaker, being careful to keep the mouth of the bottle below the surface of the water. The white powder formed by the burning phosphorus floats in the air in the bottle but is gradually dissolved in the water.

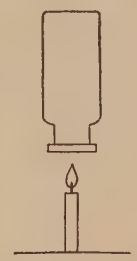

FIG. 32. Collecting the hot gases evolved by a burning candle

Leave the bottle in position until the powder has entirely disappeared. Note that the water has risen in the bottle. (*h*) How do you account for this fact? (*i*) Is your conclusion in

accord with the results obtained by burning iron in air? (*j*) Suppose it were possible for you to collect and weigh the white powder formed by the burning phosphorus, how should you expect its weight to compare with that of the original phosphorus?

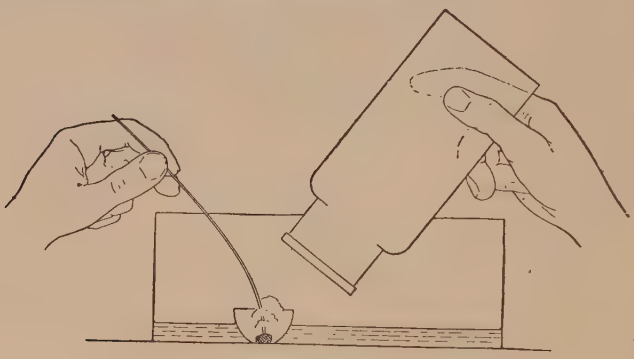

FIG. 33. Burning phosphorus in air

Remove the crucible from the water and place it on a ring stand in the hood (Fig. 30). Now heat the crucible very gently (*care*) so as to ignite any unburned phosphorus. When cool, the crucible may then be cleaned and put away.

EXERCISE 9

THE PREPARATION AND PROPERTIES OF HYDROGEN

Apparatus. Test tube; apparatus, bottles, and trough as shown in Fig. 35 (the bottles are 250-cc. wide-mouthed); beaker; stirring-rod; 60-cc. bottle; forceps; 4 glass plates; (evaporating-dish; funnel).

Materials. Bit of sodium (size of a small pea); wooden splints; 10 g. mossy zinc; 1 cc. copper sulfate solution (R.S.); 1 cc. kerosene (R.S.); sulfuric acid; (filter paper).

1. Preparation from water. Fill a test tube with water and invert it in a beaker of water. Wrap a piece of sodium in a bit of filter paper previously moistened with kerosene. The sodium, when wrapped in the paper, must be small enough to enter the tube easily. Raise the inverted test tube until its mouth dips just below the surface of the water in the beaker and, grasping

the sodium and paper with the forceps, quickly insert the sodium into the test tube (Fig. 34). Stand at arm's length, as a slight explosion sometimes occurs. Notice that the sodium decomposes the water, liberating a gas which is caught in the tube. When the action has ceased, place your thumb tightly over the mouth of the tube to prevent the gas from escaping, and bring the tube to an upright position. After lighting a splint, remove the thumb from the tube and quickly bring the flame to the mouth of the tube. (*a*) Does the gas act like oxygen? (*b*) What is the source of the gas? (*c*) Record any other methods that may be employed for obtaining it from the same source (consult text).

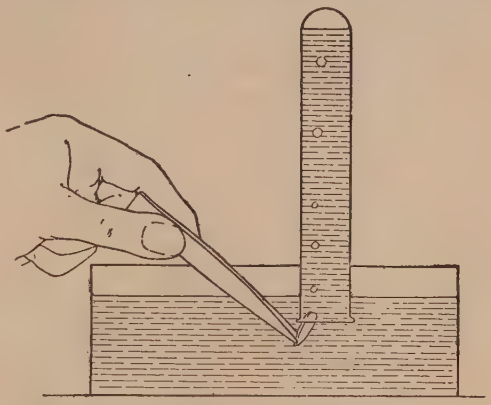

FIG. 34. Preparing hydrogen by the action of sodium on water

2. Preparation from acids (usual laboratory method). Prepare a hydrogen generator according to Fig. 35. *D* represents a wide-mouthed bottle of about 250-cc. capacity. The gas-delivery tube *B*, *C* is the same as that used in the preparation of oxygen (Fig. 28). A rubber stopper for the bottle *D* is preferable, although a good cork stopper will do provided it can be made air-tight. The funnel tube *A* must extend nearly to the bottom of the bottle. Put 10 g. of the mossy zinc into the bottle *D*. Pure sulfuric acid will not react readily with pure zinc; hence it is advisable to add to the zinc from 8 to 10 drops of a solution of copper sulfate, which will start the reaction. Now pour through the funnel tube into the bottle just enough water to cover the zinc.

Prove that the apparatus is air-tight by blowing into the delivery tube *C* until the water is forced nearly to the top of the funnel tube; then quickly close the rubber tube either by tightly pinching it or by placing the tongue firmly against its end. If the apparatus is air-tight, the water in the funnel tube will not fall.

Now prepare some dilute sulfuric acid by pouring *slowly*, a few drops at a time, 15 cc. of concentrated acid into a beaker containing 50 cc. of water. *Stir the water with a glass rod while*

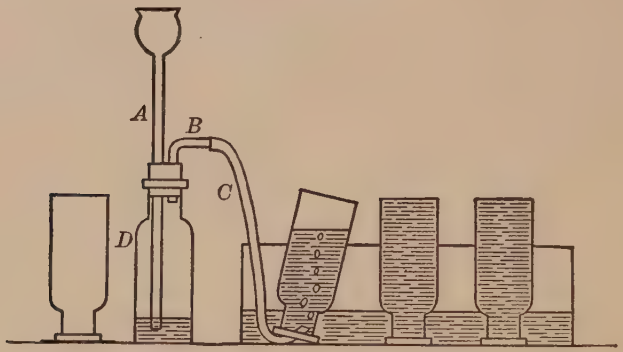

FIG. 35. Apparatus used for preparing hydrogen by the action of sulfuric acid on zinc

the acid is being added. Notice that the acid is poured into the water—never the reverse. (*Sulfuric acid mixes with water, evolving a great deal of heat. Unless the above precautions are taken, the heat may be so intense as to cause the liquid to spatter, and the hot acid in contact with the skin causes a serious burn.*)

Cool the dilute acid and pour a few drops of it through the funnel tube (*A*). Hydrogen is evolved as soon as the acid comes in contact with the zinc. Enough of the acid must be added from time to time to cause a gentle and continuous evolution of the gas. An excess of the acid should be avoided, however, or the action will become too violent, and a large quantity of zinc will have to be added at the close of the exercise.

It is evident that the gas which passes over first is a mixture of hydrogen and air. *The student must remember that such a confined mixture of hydrogen and air explodes with great violence if ignited*; hence see that the end of the delivery tube is not brought near any flame. Determine when the hydrogen is free from air by repeatedly collecting a test tube full of gas and igniting it, holding the tube mouth downward. (*d*) Why downward? If pure, the gas burns quietly; otherwise there is a slight explosion. After all the air has been expelled from the generator, collect four bottles (250-cc. wide-mouthed) of the gas; then cover the mouth of each bottle with a glass plate and place the bottles *mouth downward* on your desk.

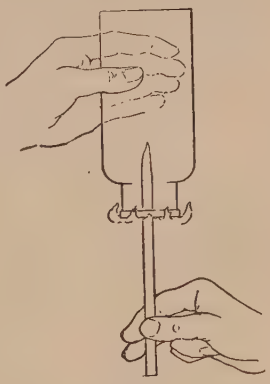

Fig. 36. Thrusting a lighted splint into hydrogen

(*e*) What is the source of the hydrogen? (*f*) What is the use of the zinc? (*g*) Why is *mossy* zinc used?

Remove the cork from the generator, add a few more pieces of zinc, and set aside for use in 4, below.

3. Properties and chemical conduct of hydrogen. Thrust a lighted splint into a bottle of the gas held mouth downward (Fig. 36). Slowly withdraw the splint and again thrust it into the gas. (*h*) Describe the results. (*i*) What do they prove?

Fill a small (60-cc.) wide-mouthed bottle or test tube one third full of water and invert it in a pneumatic trough. Displace the remaining water with hydrogen from one of the bottles. (*j*) What does the bottle now contain? Withdraw it from the water and, holding it at arm's length, quickly bring it, mouth downward, over a flame. (*k*) What do the results prove?

Uncover a bottle (mouth upward) of the gas. After a minute, test for the presence of hydrogen with a lighted splint. Repeat, holding the bottle mouth downward. (*l*) What do the results

show? (*m*) Summarize briefly the properties and chemical conduct of hydrogen.

4. Without removing the fragments of any undissolved zinc, pour the liquid left in the generator *D* in experiment 2 into an evaporating-dish and boil gently on a ring-stand support. As soon as white crusts begin to form on the side of the dish, just above the liquid, filter the hot liquid into a beaker and set the filtrate aside to cool. (*n*) How does the product which separates from the filtrate compare in properties with the original zinc? (*o*) What is the product (consult text)? (*p*) How does it differ in composition from sulfuric acid?

EXERCISE 10

THE COMBUSTION OF HYDROGEN

Apparatus. Hydrogen generator *A* (Fig. 37) attached by rubber tubing *E* to a drying-tube *B*. This tube is filled with granulated calcium chloride, held in place by loose plugs of cotton placed at each end of the tube. *C* is a glass tube drawn to a jet; rubber tube *C* (Fig. 35).

Materials. Granulated calcium chloride sufficient to fill tube *B* (Fig. 37); cotton sufficient to use in tube *B*; 8 g. zinc; 1 cc. copper sulfate solution (R.S.); dilute sulfuric acid for preparing hydrogen (see Exercise 9); picture-frame wire 10 cm. long; bit of charcoal.

1. Charge the generator *A* (Fig. 37) with 6 or 8 g. of zinc, add the solution of copper sulfate, cover with water, and add dilute sulfuric acid as in 2, Exercise 9, taking all the precautions noted in the exercise. The hydrogen passes from the bottle through the tube *E*, is freed from all moisture in passing over the calcium chloride in the tube *B*, and escapes through the tube *C*. Slip a piece of rubber tubing over the tube *C*, collect samples of the gas in test tubes over water, and test with a flame to see whether the hydrogen is free from air. *After all the air has been expelled, wrap a towel carefully about the generator*, remove the rubber tubing, and cautiously ignite the hydrogen escaping at the end of tube *C*. The flame is nearly invisible and is very hot. Test the heat of the flame by holding in it different objects, such as a splint, a piece of picture-frame wire, a bit of charcoal.

Finally, hold over the flame a cold, dry beaker or bottle and note the liquid (water) deposited on the sides of the vessel.

(*a*) Summarize briefly all the facts brought out in the above experiment.

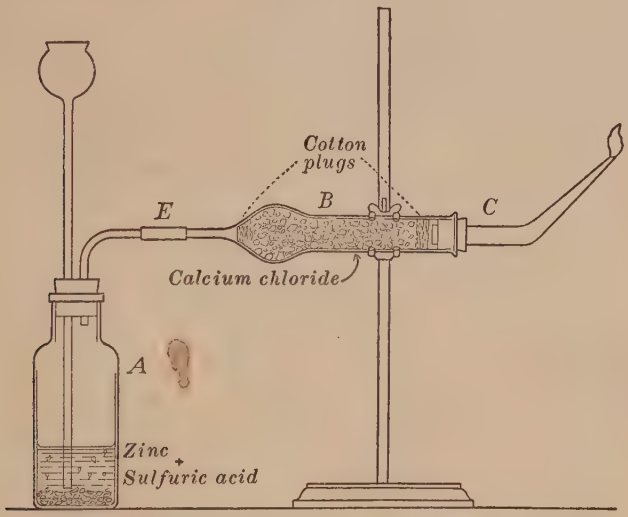

FIG. 37. Apparatus for preparing and burning hydrogen

EXERCISE 11

OXIDATION AND REDUCTION

Apparatus. Hydrogen generator and tubes as shown in Fig. 38 (*A* is the hydrogen generator, *B* is a drying-tube filled with calcium chloride, *C* is a straight glass tube, and *D* is a hard-glass test tube); apparatus used in preparing oxygen (Fig. 28).

Materials. 2 g. copper oxide; calcium chloride sufficient to fill the drying-tube *B*; 8 g. zinc; dilute sulfuric acid for preparing hydrogen (see Exercise 9); copper sulfate solution (R. S.); 4 g. potassium chlorate; 2 g. manganese dioxide.

1. Remove the test tube *D* (Fig. 38), introduce into it 1 or 2 g. of copper oxide, and set it aside.

Now generate hydrogen in *A* as in Exercise 10. *After all the air has been expelled from the apparatus* (test samples to be certain), wrap the generator in a towel as described in Exercise 10. Then slip the test tube *D* over the exit tube *C*, as shown in Fig. 38, and cautiously heat the copper oxide in *D* to redness, being careful to keep the flame away from the mouth of the tube. Note the condensation of moisture in the cold portions of

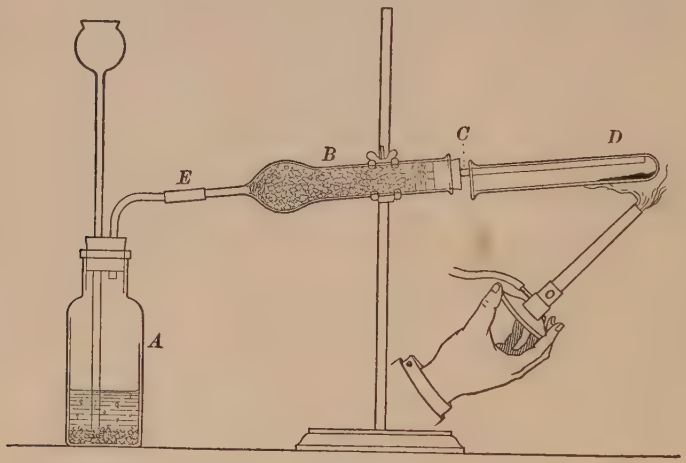

FIG. 38. Reducing copper oxide by hydrogen

the tube. (*a*) Is there any visible evidence of change in the copper oxide? (*b*) Explain. (*c*) What is the object of the calcium chloride in tube *B*?

2. Disconnect the bottle *A* at *E* from the rest of the apparatus and blow air through the tube *B* to remove the hydrogen present. Now connect the apparatus used in preparation of oxygen (Fig. 28) at *E* to the tubes *B*, *C*, and *D*. Generate oxygen as in Exercise 7 and conduct a slow current of the gas through *B*, *C*, and *D*, and at the same time gently heat the residue in *D*. (*d*) Record your results.

3. (*e*) Define the terms *reduction* and *oxidation* and give an example of each from the above experiment.

EXERCISE 12

THE MEASUREMENT OF GAS VOLUMES

Apparatus. Graduated tube and cylinder as shown in Fig. 39; thermometer.

1. (It is suggested that the instructor arrange one or more pieces of apparatus as shown in Fig. 39. The students will then take the readings and solve the problems.)

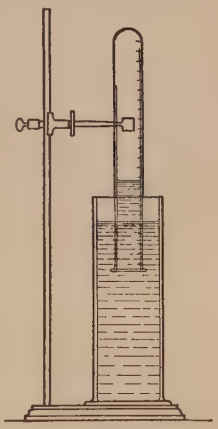

FIG. 39. Measuring a volume of gas over water

Partially fill a graduated tube with water and invert it in a cylinder (or other vessel) of water as shown in Fig. 39. Let the water stand until its temperature is the same as that of the air. Adjust the tube until the level of the liquid inside and outside of the tube is the same; then take the reading of the volume of the air in the tube. Note the temperature of the air (place a thermometer by the side of the tube) and likewise the pressure of the atmosphere as indicated by the barometer. Insert these values in their appropriate places in the following table; then calculate the volume which the air would occupy under standard conditions.

Volume of air in tube	cc.
Temperature of air	°
Barometric pressure	mm.
Vapor pressure of water at temperature of air (see Appendix)	mm.
Effective pressure on air in tube (barometric pressure less vapor pressure)	mm.
Volume which the air would occupy under standard conditions (calculate)	cc.

EXERCISE 13

BOILING POINTS AND FREEZING POINTS; AMORPHOUS AND CRYSTALLINE SUBSTANCES

Apparatus. Flask (250-cc.) fitted with a two-hole cork, through which pass the thermometer *B* and the glass tube *C*, as shown in Fig. 40; test tube; large beaker; small beaker; stirring-rod; watch glass.

Materials. 300 g. ice; 100 g. common salt; 3 g. powdered alum; 5 g. powdered sulfur; 1 cc. chloroform.

1. Boiling points. Fill the flask *A* (Fig. 40) about one third full of water. Lower the thermometer until the bulb is immersed in the liquid. Heat the water to boiling and note the temperature after it has become constant. Apply a greater heat and note the effect upon the temperature.

Raise the thermometer until the bulb no longer touches the boiling water but is simply exposed to the steam. Again note the temperature. (*a*) Record your general conclusions.

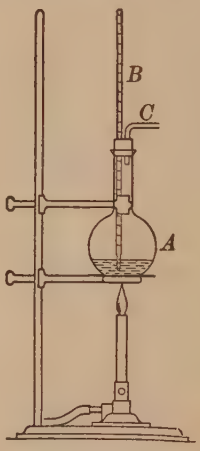

FIG. 40. Apparatus for determining the boiling point of a liquid

2. Freezing points. Place some pieces of ice in a small beaker. Introduce 15 or 20 cc. of water and stir the mixture gently with the thermometer. (*b*) Continue until the temperature is constant, and record the temperature.

Half fill a test tube with pure water and dip the tube into a freezing-mixture made by mixing 3 parts of powdered ice with 1 part of common salt. Stir the water in the test tube from time to time with a thermometer and note the temperature at which it begins to freeze. (*c*) Compare the freezing point of water with the melting point of ice.

3. Add powdered sulfur to a test tube until the tube is about one third filled; then heat the tube gently until the sulfur is melted. Set the tube aside until the liquid has again solidified. (*d*) Is the solid in a crystalline form?

4. Place about 3 g. of powdered alum in a small beaker or evaporating-dish and dissolve the solid in as little water as possible, at room temperature. Hasten the solution by stirring the mixture with a glass rod. In the same way make a solution of common salt. Set the solutions aside (do not cover the beaker or dish) until the next laboratory period. (*e*) Note the form in which each solid is deposited, illustrating, if possible, by a drawing. (*f*) Mention two common methods used for obtaining substances in crystalline state.

5. Pour about 1 cc. of chloroform on a watch glass and blow across the surface of the liquid. (*g*) Account for its rapid evaporation. (*h*) Note and account for any change in the temperature of the glass.

EXERCISE 14

THE PROCESS OF DISTILLATION; THE COMPOSITION OF WATER

Apparatus. Flask (250-cc.); Liebig condenser (*B*) and connections as shown in Fig. 41, or apparatus shown in Fig. 42; wire gauze 12 cm. square; watch glass; glass rod.

Materials. 10 cc. alcohol.

1. Distillation of water. (Two or more students may work together if the apparatus is not available for each.) Connect a Liebig condenser *B* with a 250-cc. flask *A*, as represented in Fig. 41. The flask is set on a piece of wire gauze supported by the iron ring attached to the ring stand. The tube *C* is connected with the water pipe by means of rubber tubing, and a current of cold water is conducted through the outer tube of the condenser, escaping through the tube *D*. (*a*) Why is cold water admitted at *C* rather than at *D*?

Fill the flask one fourth full of hydrant or well water and boil the water until 25 cc. or more of liquid has collected in the *receiver E*. If condensers are not available, the apparatus shown in Fig. 42 may be used. The steam from the boiling water in flask *A* is condensed by conducting it through *B* into the test tube *C*, which is kept cold by ice water in a beaker or bottle *D*.

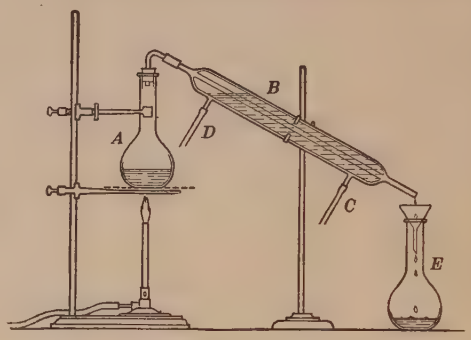

FIG. 41. Distilling water in the laboratory

(*b*) Compare the *distillate* (distilled water) with the hydrant water in appearance and in taste.

Place 4 or 5 drops of the distilled water on a watch glass and evaporate, holding the watch glass 10 or 15 cm. above the tip of the flame. (*c*) Is there any residue? (*d*) Repeat, using hydrant water. (*e*) Why is distilled water used in the laboratory?

Repeat the distillation, using a sample of muddy water in *A*, and (*f*) record the results.

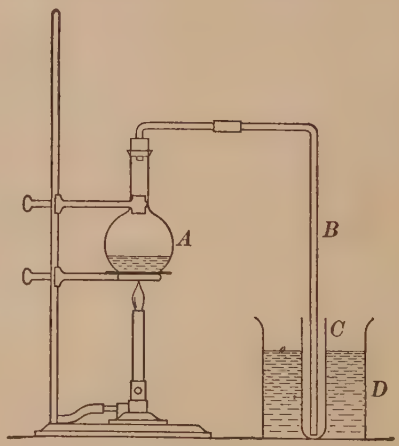

FIG. 42. A simple apparatus for distilling small volumes of water

2. Fractional distillation. Dip the end of a glass rod into alcohol and at once bring it into contact with a flame. (*g*) Is the alcohol inflammable? Now distill (Fig. 41) a mixture of 10 cc.

of alcohol (boiling point 78.3°) and 30 cc. of water. Collect the first 1 or 2 cc. of the distillate in an evaporating-dish and test with a flame. In the same way test successive portions of the distillate. (*h*) Does there seem to be a partial separation of the two liquids? In this way a mixture of liquids boiling at different temperatures may generally be separated more or less perfectly. The process is termed *fractional distillation.*

3. **The composition of water.** Recall the experiments in Exercise 10, also in Exercise 11. (*i*) What do these experiments show in regard to the elements present in water? (*j*) Can you suggest any modification of Exercise 11 that would enable you to determine the relative weights of the elements present in water?

EXERCISE 15

THE PROPERTIES OF HYDROGEN PEROXIDE

Apparatus. 3 test tubes; funnel; small beaker.

Materials. 10 cc. hydrogen peroxide solution; 1 g. manganese dioxide; wooden splint; filter paper; 1 cc. of starch solution (R. S.); small crystal of potassium iodide.

1. (*a*) What is the strength of the hydrogen peroxide solution sold by druggists (consult text)? Pour 3 cc. of the solution into a large test tube and add 1 or 2 g. of finely powered manganese dioxide. (*b*) Test the gas evolved with a glowing splint.

Filter the mixture remaining in the tube. The solid is the unchanged manganese dioxide. (*c*) In what other experiment has manganese dioxide brought about a change without apparently undergoing any change itself?

2. To 1 cc. of starch solution add 4 cc. of water. Dissolve in this a small crystal of potassium iodide and then add a few drops of a solution of hydrogen peroxide. The peroxide liberates the iodine from the potassium iodide, and the resulting free iodine colors the starch paste. (*d*) Note the color.

EXERCISE 16

THE PREPARATION AND PROPERTIES OF NITROGEN

Apparatus. Apparatus shown in Fig. 43: *A* is the *hard*-glass tube included in your outfit and is supported on a ring stand. It contains a spiral of fine copper gauze (from 60 to 70 mesh), made by rolling up a piece of gauze about 10 × 12 cm. The gauze must fit the tube loosely so that it can be pushed easily to the middle of the tube. *B* is a 1-liter (or larger) bottle (the ordinary 5-pint acid bottle serves the purpose well). The glass tube *D* extends to the bottom of the bottle *B* and is connected to a water tap by means of a rubber tube; 2 wide-mouthed bottles (250-cc.); pneumatic trough; wing-top burner (hydrogen generator).

Materials. Pieces of copper gauze (from 60 to 70 mesh) 10 × 12 cm.; splint; (zinc and dilute sulfuric acid for generating hydrogen).

1. Preparation of nitrogen from the air. When air is passed through a tube containing hot copper, the oxygen in the air

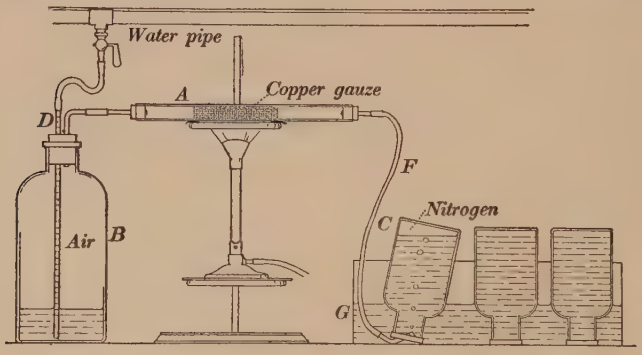

FIG. 43. Preparing nitrogen by passing air over hot copper

combines with the copper, leaving the nitrogen. This gives a simple method for preparing nitrogen. Proceed as follows: Arrange the apparatus as shown in Fig. 43 and see that the parts are all well connected. The bottle *B* is empty at the beginning, while the wide-mouthed bottles inverted in *G* are completely

filled with water. Now heat the copper spiral in the tube, using a wing-top burner, until the gauze is dull red; this will require about five minutes. Then continue the heating, and at the same time force a *slow* current of water into *B* at a rate not to exceed 1 l. in twenty minutes. The water entering *B* forces the air in the bottle over the hot copper, which unites with the oxygen, while the nitrogen passes on and is collected in the bottles in *G*. When two of these bottles are filled, disconnect the apparatus and withdraw the heat. (*a*) Why disconnect the apparatus before withdrawing the heat?

Remove the bottles from *G* as in the preparation of oxygen and hydrogen, keeping each covered with a glass plate. Since the copper removes only the oxygen, the nitrogen prepared by this method will contain small percentages of the rare elements present in the atmosphere. Since they are present in such small amounts (about 1 per cent of the original air) and resemble nitrogen so closely, they do not materially affect the properties of the nitrogen.

2. **Properties and chemical conduct of nitrogen.** (*b*) Note the color, odor, taste, and solubility in water of nitrogen. Insert a burning splint into a bottle of the gas. (*c*) Is nitrogen combustible? (*d*) Is it a supporter of combustion? (*e*) Summarize briefly the properties of nitrogen.

3. Since the copper spiral is partly oxidized to copper oxide in the experiment, and since you will want the copper spiral in a later experiment, it is advisable, if time is available, to reduce the spiral back to copper. To do this, connect your hydrogen generator to the tube *A* containing the spiral, and *after all the air is expelled from the apparatus* (test as in Exercise 9) gently heat the tube containing the gauze, while passing a slow current of hydrogen through the tube. The gauze is soon reduced, as shown by the change in color. Leave the copper in the tube until cold. (*f*) Why until cold?

EXERCISE 17

THE MOLECULAR WEIGHT OF OXYGEN; THE PERCENTAGE OF OXYGEN IN POTASSIUM CHLORATE (Optional)

Apparatus. Apparatus shown in Fig. 44 (*A*, *B*, *C*, is the apparatus used in the preparation of oxygen, Fig. 28. Care must be taken that the glass tube and cork fit the test tube *air-tight*, so that any gas generated in the tube can escape only through the exit tube *C*. *D* is an ordinary 5-pint acid bottle, which must be completely filled with water and inverted in the water in the trough); cork to fit the bottle *D*; porcelain crucible; iron wire 10 or 12 cm. long; graduated cylinder.

Materials. 3 g. powdered manganese dioxide; 7 g. potassium chlorate.

1. Principle involved. The molecular weight of a gaseous element or compound is determined by obtaining the weight of a definite volume of the gas. Since the direct weighing of a gas is not easily carried out, it is better to adopt some method which avoids the direct weighing of the gas. In the case of oxygen this is readily done by selecting a solid compound of oxygen (such as potassium chlorate) which will evolve the gas when heated. The difference in the weights of the compound before and after heating gives the weight of the oxygen evolved. By collecting the oxygen evolved it is easy to determine its volume. We have then the volume occupied by a definite weight of oxygen, and from these results we can calculate the weight of 22.4 l. of the gas. Proceed as follows:

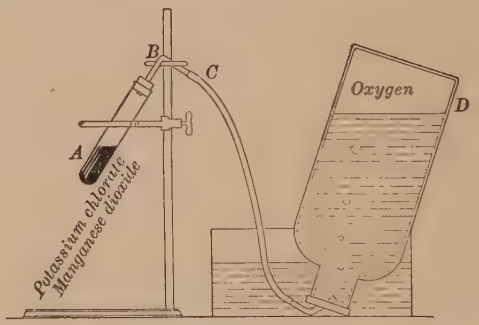

Fig. 44. Determining the volume of a definite weight of oxygen

2. Preparation of the oxygen. In your porcelain crucible place about 3 g. of finely powdered manganese dioxide and heat it strongly in the Bunsen flame for about five minutes, stirring it from time to time with a clean iron wire or end of file; then withdraw the flame and allow the powder to cool.

Thoroughly clean and dry your hard-glass test tube (*A*, Fig. 44) and weigh it accurately to 0.01 g. Record the weight in the table below. Weigh out roughly between 6 and 7 g. of potassium chlorate and introduce this into the tube (Fig. 26), taking care that none of the chlorate adheres to the sides of the tube. Accurately weigh (to 0.01 g.) the tube and contents as before, recording the weight in the table. Now introduce into the tube the manganese dioxide contained in the crucible, and again weigh to 0.01 g. and record the weight in the table. Mix the two compounds in the bottom of the tube by gently rotating the tube and by tapping it against your hand. Connect the tube as shown in Fig. 44 and arrange the exit tube so that all gas evolved will be collected in the bottle *D*, which must be completely filled with water. Now hold the burner in your hand and heat the mixture gently. Oxygen is evolved and is collected in the bottle. Regulate the flame so that the bubbles of the gas *can easily be counted.* Continue the heating *until the gas is no longer evolved.* Remove the cork from the test tube and carefully set the tube aside to cool.

Now adjust the bottle containing the oxygen so that the level of the water inside and outside the bottle is the same (it may be necessary to add more water to the trough); then, while holding it in this position, push a cork tightly into the bottle and quickly bring the bottle to an upright position on the desk. Next, take the temperature of the water in the trough and the reading of the barometer and record the figures in the table. If time is not available the bottle containing the oxygen (and some water), and the tube *A*, in which the oxygen was prepared, can be set aside and the experiment concluded later.

3. **Determination of the weight of 22.4 l. of oxygen.** Weigh to 0.01 g. the test tube and residue left after preparing the oxygen, recording the weights in the table. The loss in the weight of the tube and contents on heating equals the weight of the oxygen evolved. Now measure the volume of the oxygen in the bottle by adding water from the graduated cylinder and record the volume. The volume occupied by the oxygen is under laboratory conditions. Reduce this to standard conditions and record the result in the table. You now have the volume under standard conditions occupied by a known weight of oxygen. From these results calculate the weight of 1 l. of oxygen under standard conditions and from this, the weight of 22.4 l. of oxygen. Record the weights in the table.

4. Calculation of percentage of oxygen in potassium chlorate. All the oxygen evolved in the above experiment comes from the potassium chlorate; moreover, if the experiment is carried out as directed, all the oxygen in the potassium chlorate is evolved, since the mixture is heated until the gas ceases to be given off. From the weight of the potassium chlorate taken and the weight of the oxygen evolved, calculate (see table) the percentage of oxygen in the potassium chlorate and record your results in the table.

Weight of test tube *A*	g.
Weight of test tube plus potassium chlorate . . .	g.
Weight of potassium chlorate taken (equals the difference of the two weights given above) . . .	g.
Weight of the test tube plus potassium chlorate plus manganese dioxide	g.
Weight of the test tube and residue left after heating	g.
Weight of oxygen evolved (equals loss in weight of tube and contents on heating)	g.
Temperature of water in pneumatic trough . . .	°
Reading of barometer	mm.
Volume of oxygen in bottle *D*	cc.
Volume of this oxygen under standard conditions (calculate)	cc.

Weight of 1 l. of oxygen under standard conditions (calculate) . . . . . g.

Weight of 22.4 l of oxygen under standard conditions (calculate). (The number expressing the weight of 22.4 l. of oxygen is the same as the number expressing the molecular weight of the gas.) . . g.

Percentage of oxygen in potassium chlorate (calculate from the following formula):

$$\frac{\text{Weight of oxygen evolved} \times 100}{\text{Weight of potassium chlorate taken}} = \left\{ \begin{array}{l} \text{percentage of oxygen in} \\ \text{potassium chlorate} \end{array} \right. \%$$

	Theoretical	Results Obtained	Average of Results Obtained by Class
Molecular weight of oxygen . .	32		
Percentage of oxygen in potassium chlorate	39.2		

(*a*) What are the sources of error in the above experiment?

EXERCISE 18

THE DETERMINATION OF THE FORMULA OF COPPER OXIDE (Optional)

Apparatus. Apparatus shown in Fig. 45: *A* is the hard-glass tube used in Exercise 16, *B* is a hydrogen generator, and *C* is a 250-cc. wide-mouthed bottle filled with water and inverted in a pneumatic trough. Porcelain boat about 8 cm. long and 6 or 7 mm. wide.

Materials. 1 or 2 g. black *finely powdered* copper oxide; zinc and dilute sulfuric acid for preparing hydrogen as in Exercise 10.

1. Thoroughly clean and dry the porcelain boat and weigh it to 0.01 g., recording the weight in the table below. Nearly fill the boat with the copper oxide and again weigh to 0.01 g. and record the weight. Carefully introduce the boat and copper oxide into the glass tube *A* (Fig. 45) and connect the apparatus (*air-tight*) as shown in the figure. Now generate hydrogen in bottle *B* as directed in Exercise 10. When the apparatus is filled with hydrogen (*note all precautions given in Exercise 10*), carefully heat the porcelain boat in the tube with a wing-top burner, maintaining all the time a slow current of hydrogen

through the tube. Gradually increase the heat and continue the operation for about twenty minutes, collecting any excess hydrogen in the bottle *C*. The hydrogen reduces the copper oxide to copper as in Exercise 11. Then disconnect the hydrogen generator, withdraw the heat, and allow the boat to remain

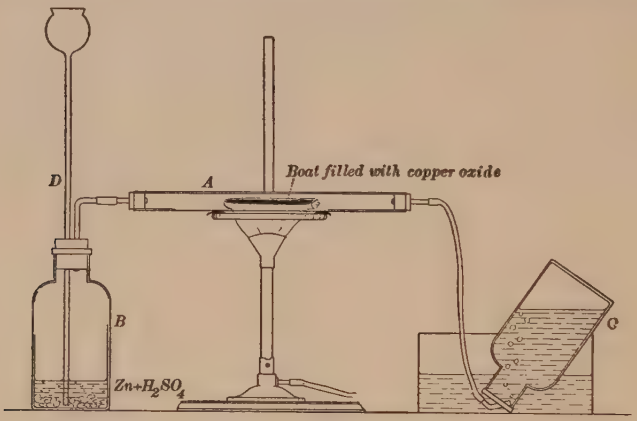

FIG. 45. Apparatus for determining the formula of copper oxide

in the tube until cold. Carefully remove the boat and contents, and weigh to 0.01 g., recording the weights in the table. (*a*) From your results calculate the percentage and composition of copper oxide; also its simplest formula as directed in the following table:

Weight of boat	g.
Weight of boat plus copper oxide	g.
Weight of copper oxide taken (calculate)	g.
Weight of boat plus residue (copper) left after heating	g.
Weight of copper left in boat (calculate)	g.
Percentage of copper in copper oxide (calculate from formula):	
$\frac{\text{Weight of copper left in boat}}{\text{Weight of copper oxide taken}} \times 100$	%
Percentage of oxygen in copper oxide (100 % − % of copper in copper oxide)	%
Simplest formula (calculate from percentage composition)	

EXERCISE 19

SOME FURTHER MANIPULATIONS

Apparatus. File; forceps; medicine dropper; (500-cc. flask and cork to fit).

Materials. 25 cm. glass tubing; (50 cm. glass tubing; 5 cm. soft rubber tubing).

1. Devices for transferring liquids drop by drop. An ordinary medicine dropper (obtainable at any drug store) is useful for transferring a few drops of a liquid from one vessel to another (Fig. 46). A glass tube made as described below is often of more service.

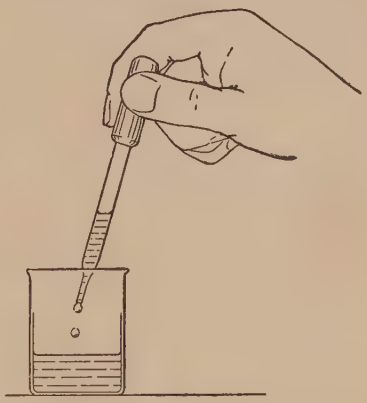

FIG. 46. Transferring a liquid, drop by drop, by the use of a medicine dropper

Select a piece of glass tubing about 25 cm. long. Heat the glass tube, about 5 cm. from one end, in the Bunsen flame until the walls of the heated portion thicken and the size of the bore diminishes (*A*, Fig. 47). The tube must be constantly rotated to prevent the softened portion from sagging.

FIG. 47. Making a pipette

Now quickly remove the tube from the flame and, holding it in a vertical position, gently pull the ends apart until the bore is of the desired size (*B*, Fig. 47). The tube is then cut at *B*,

and both ends of the longer piece are rounded by heating in the flame as directed in Exercise 3 (Fig. 16).

To transfer a liquid from one vessel into another, dip the small end of the tube into the liquid to be transferred and, placing the other end of the tube in the mouth, gently suck the liquid up in the tube until partially filled. *Great care must be taken to keep the end of the tube well below the surface of the liquid while it is being drawn up into the tube.* Moreover, *the tube should never be filled more than about two thirds; otherwise the liquid may be drawn into the mouth and serious results follow.* When the tube is partially filled, remove the end from your mouth and quickly press your forefinger firmly on the upper end of the tube. The liquid in the tube may now be transferred to another vessel by bringing the small end of the tube over the mouth of the vessel and then slightly decreasing the pressure of the finger on the tube, as shown in Fig. 48. Such a tube as the one described above is known as a *pipette.*

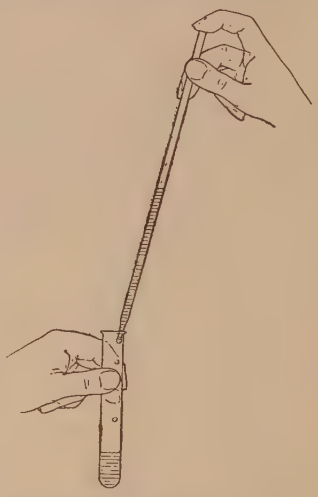

FIG. 48. Transferring a liquid, drop by drop, by means of a pipette

A pipette like the above, but graduated in cubic centimeters, serves the same purpose but has an additional advantage in that it enables one to tell just how much of the liquid is transferred, in case it is so desired. In using such a pipette, the liquid is drawn up a little above the zero point. The finger is then quickly placed against the end of the pipette, and the liquid allowed to flow out until the level of the liquid is just at the zero mark, when the flow is stopped by increasing the pressure of the finger against the end of the pipette. The

liquid may then be allowed to drop (or flow) into any vessel until the desired amount is transferred. The flow is stopped by increasing the pressure of the finger on the end of the pipette.

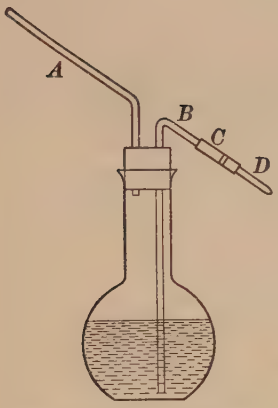

FIG. 49. A wash-bottle

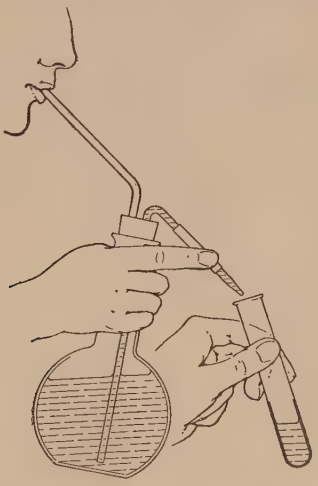

FIG. 50. Method of using a wash-bottle

2. While not essential, the student will find a so-called wash-bottle very convenient for adding water to a test tube or beaker. A 500-cc. flask or bottle is a convenient size. The apparatus is shown in Fig. 49. *A*, *B*, and *D* are glass tubes; the tube *D* is drawn out to a small bore (about as large as the lead in a lead pencil). The tubes *B* and *D* are joined by the rubber tube *C*. The bottle is filled with water, and a stream is projected through *D* by blowing in the tube *A*. Fig. 50 shows the method of using the bottle.

EXERCISE 20

THE FORMATION OF CHARCOAL AND COKE

Apparatus. Hard-glass test tube *A* connected as shown in Fig. 51; *B* is a test tube fitted with a glass tube *C* drawn to a jet; large beaker or bottle.

Materials. Small pieces of hard wood (sawdust will do) sufficient to half fill the hard-glass test tube; pieces of soft coal sufficient to half fill the test tube.

1. Half fill the tube *A* (Fig. 51) with hard-wood splints or sawdust and connect it as shown in the figure. The tube *B* is kept cool by ice water in the beaker (or bottle). Heat the wood in *A*, gradually increasing the heat until no further change takes place. (*a*) During the heating, test the gas escaping from the jet *C* to determine whether it will burn.

When the tube *A* is cool, remove the residue and note its properties. (*b*) Is it combustible? (*c*) This experiment illustrates the manufacture of what substance?

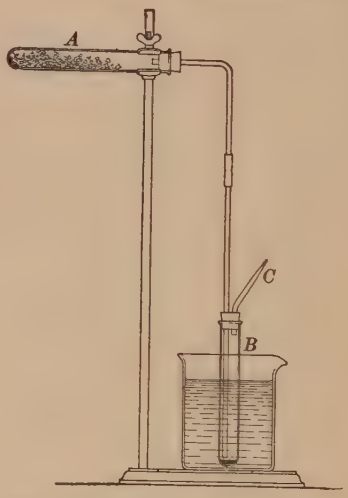

FIG. 51. Apparatus for heating coal and wood in the absence of air and for collecting the products formed

(*d*) Note the odor and appearance of the liquid condensed in *B*. (*e*) Name two important compounds prepared commercially from the liquid obtained by heating hard wood in the absence of air (consult text). (*f*) What name is applied to the process which takes place when any combustible substance, such as wood or coal, is heated in the absence of air?

2. Repeat the experiment, substituting small pieces of *soft* coal for the wood. (*g*) Describe the results. (*h*) What is left in the tube *A*? The black liquid condensed in *B* is known as *coal tar* (consult text). (*i*) Note its appearance and odor.

EXERCISE 21

A FURTHER STUDY OF CARBON

Apparatus. Hard-glass test tube; porcelain dish; 250-cc. flask; funnel.

Materials. 2 g. sugar; test tube one fourth full of boneblack; 1 cc. litmus solution (R. S.); filter paper; 2 pieces of charcoal the size of a pea; sulfuric acid; hydrochloric acid.

1. Heat 1 or 2 g. of sugar in a test tube until no further change takes place. (*a*) Note the results. (*b*) What is the residue?

2. Bring a cold porcelain dish into a small luminous Bunsen flame. (*c*) Note the deposit. (*d*) What is this form of carbon called? (*e*) In what other forms does carbon exist?

3. Fill a test tube about one fourth full of boneblack and transfer it to a small flask; then pour over the boneblack about 50 cc. of water to which has been added a few drops of a solution of litmus or indigo. Thoroughly mix the contents of the flask; then heat gently for a few minutes and filter. If the filtrate is not decolorized, repeat the process, using more boneblack. (*f*) What is the composition of boneblack? (*g*) By what other name is it known? (*h*) What use does this experiment suggest for it?

4. Place a piece of charcoal in a test tube and cover it with sulfuric acid. Does any action seem to take place? Repeat, using hydrochloric acid. (*i*) Is carbon an active element at ordinary temperatures? (*j*) How does the charring of wood preserve it?

EXERCISE 22

THE FORMATION OF CARBON DIOXIDE FROM CARBON AND OXYGEN

Apparatus. Apparatus as shown in Fig. 52. Tube *A* and bottle *B* are the same as described in Fig. 43, Exercise 16. *E* is a glass tube, and *C* is a small bottle containing about 10 cc. of limewater; wing-top burner.

Materials. 15 cc. limewater; 5 or 6 bits of charcoal the size of a bean.

1. Test for carbon dioxide. Repeat the first part of experiment 2, Exercise 8. Limewater is a solution of ordinary slaked lime in water. Slaked lime is a white solid compound of the formula CaO_2H_2 and is known in chemistry as *calcium hydroxide*. The change in the limewater produced by carbon dioxide is due to the fact that the carbon dioxide acts upon the calcium hydroxide in the solution to form calcium carbonate ($CaCO_3$) thus:

$$CO_2 + CaO_2H_2 \longrightarrow CaCO_3 + H_2O$$

The calcium carbonate so formed is a white solid and is insoluble in water, and hence separates, as fast as formed, thus causing the limewater to become cloudy. *This reaction serves as a good test for carbon dioxide.*

2. Arrange an apparatus as shown in Fig. 52. Place in the glass tube *A* 5 or 6 small pieces of charcoal and connect the apparatus as shown. Pour about 10 cc. of clear limewater into the bottle *C*. Now turn the water cock slightly so that a slow current of water flows into the bottle *B*. This forces the air in the bottle out and over the charcoal in *A* and finally up through the limewater in *C*. Continue for one or two minutes. (*a*) Is there any change in the appearance of the limewater? Turn off the water and gradually heat the charcoal in tube *A* with the wing-top burner. When the charcoal is hot continue the heating and again turn on the water so that a slow current of

air is forced over the hot charcoal. (*b*) Note again any change in the limewater. (*c*) Is carbon dioxide formed in the experiment? (*d*) What is its source? (*e*) Write the equation for its

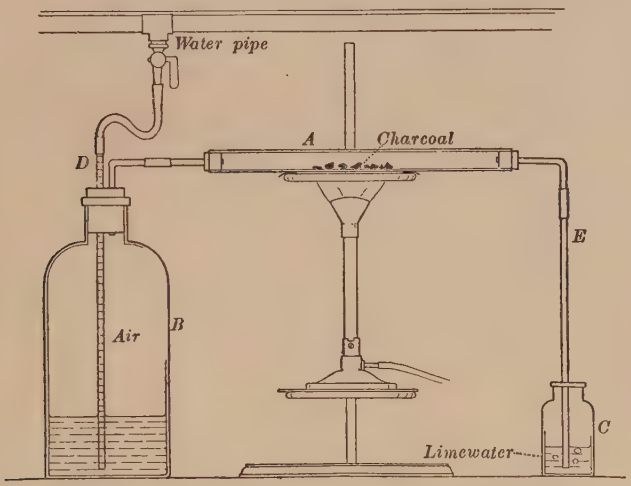

FIG. 52. The formation of carbon dioxide by the direct combination of carbon and oxygen

formation. (The percentage of carbon dioxide in the air is so small that it has no appreciable effect on the limewater during the time of the experiment.)

EXERCISE 23

A STUDY OF CARBON DIOXIDE

Apparatus. Hydrogen generator with connections, as shown in Fig. 53; three 250-cc. bottles; small beaker or test tube; hard-glass test tube; glass rod.

Materials. 5 pieces of marble (size of walnuts); hydrochloric acid; splints; limewater (R.S.); 3 g. of copper oxide and an equal bulk of powdered charcoal.

1. Usual laboratory method for preparing carbon dioxide. Place some pieces of marble in your hydrogen generator and

connect as shown in Fig. 53. Add water through the funnel tube until the marble is covered with the liquid; then add hydrochloric acid, a few drops at a time. Carbon dioxide is formed, passes out through the exit tube, and collects in the bottle *B*. Collect three bottles of the gas by displacement of air. To test when filled, hold a burning splint at the mouth of the bottle: the gas will extinguish the flame. (*a*) Why not collect the gas over water as in the case of oxygen and hydrogen?

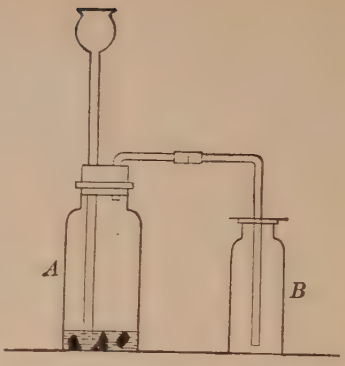

FIG. 53. The preparation of carbon dioxide by the action of hydrochloric acid on marble

2. Properties of carbon dioxide. (*b*) Note the color and odor of the gas. (*c*) Pour the gas in one of the bottles into another bottle filled with air, just as you would pour water; then test for the gas in each of the bottles with a lighted splint. (*d*) What do the results show in reference to the density of the gas as compared with air?

3. Chemical conduct of carbon dioxide. Introduce a lighted splint into a bottle of the gas. (*e*) Is the gas combustible? (*f*) Is it a supporter of combustion?

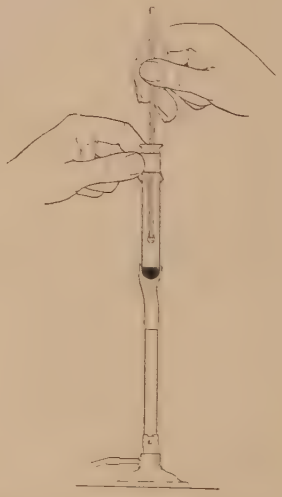

FIG. 54. Testing for carbon dioxide, using a drop of limewater

Pour 5 cc. of clear limewater into one of the bottles; then close the bottle with your hand and shake the liquid. (*g*) Explain the results.

(*h*) Mix together in a mortar 2 or 3 g. of black copper oxide and an equal bulk of powdered charcoal. Transfer to a hard-glass test tube and heat gently. The copper oxide is reduced to copper, the oxygen combining with the hot carbon to form carbon dioxide. Prove the presence of carbon dioxide in the tube as follows: Dip the end of a glass rod into limewater and withdraw it gently so that a drop of the clear liquid remains clinging to the end of the rod. Lower this carefully into the tube (Fig. 54). Any carbon dioxide present will cause the drop of liquid to become cloudy. (*i*) What kind of an agent is carbon? (*j*) Briefly summarize the properties and chemical conduct of carbon dioxide.

EXERCISE 24

THE PREPARATION AND PROPERTIES OF CARBON MONOXIDE (Optional)

Apparatus. The apparatus shown in Fig. 55: *A* is a funnel, connected by a rubber tube with a piece of glass tubing which passes through a stopper into the 250-cc. flask *B*; *C* is a small clamp; pneumatic trough; 3 wide-mouthed bottles; 3 glass plates 10 cm. square.

Materials. Sulfuric acid; 25 cc. formic acid (50 per cent); limewater (R. S.).

Precaution. *Carbon monoxide is a nearly odorless and very poisonous gas and must not be breathed. All the experiments must be performed in the hood. After the gas is generated, pour the contents of the generator flask into a sink or jar in the hood. If a good hood is not available, the exercise should be omitted.*

1. Preparation of carbon monoxide. Remove the stopper from the flask *B* (Fig. 55), pour in 15 cc. of sulfuric acid, and connect the apparatus as shown in the figure. Close the clamp *C* and partially fill the funnel *A* with the formic acid. Now open the clamp carefully so that the formic acid will enter the flask drop by drop. Allow 8 or 10 drops to flow in; then close the clamp. If the reaction does not begin (as indicated by

absence of effervescence of the liquid in the flask and escape of gas through the exit tube), heat the flask very gently until the reaction starts; then open the clamp again and admit the formic acid, a drop at a time, so as to secure a regular flow of gas from the flask. If necessary, add more formic acid to the funnel so as to keep it partially filled. Collect three bottles of the gas as shown in the figure, being careful to note which of the bottles was the first one filled. Close the clamp so as to stop further generation of gas. Slip the glass plates over the mouths of the bottles and remove the bottles from the trough, keeping them closed with the glass plates. (*a*) Write the equation for the preparation of the oxide from formic acid.

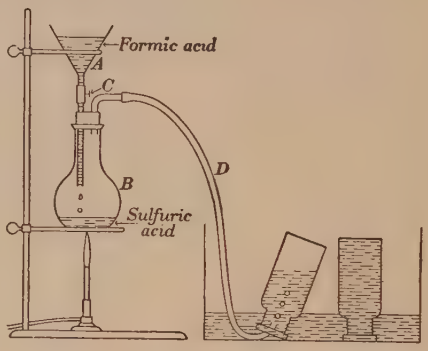

FIG. 55. Preparing carbon monoxide by the action of sulfuric acid on formic acid

2. Properties and chemical conduct of carbon monoxide. (*b*) Has the gas any color? (*c*) Is it soluble in water?

(*d*) In the first bottle filled under 1 is the gas pure carbon monoxide? Test it by introducing a burning splint into the bottle; then repeat with the second bottle. (*e*) Is the gas combustible? (*f*) Is it a supporter of combustion? Slip the glass plate from the third bottle just far enough to pour into the bottle 5 cc. of clear limewater; then quickly replace the glass plate and, holding it firmly against the mouth of the bottle, shake the contents of the bottle. Note any change in the appearance of the limewater. Now tip the bottle as far as possible without spilling the limewater; remove the glass plate and quickly ignite the gas, holding the bottle in this position so that at least a portion of the combustion product may be retained in

the bottle. When the flame dies out, at once cover the mouth of the bottle with the glass plate and shake the contents. (*g*) Record and explain the results.

3. (*h*) Summarize the properties and chemical conduct of carbon monoxide.

EXERCISE 25

THE DETERMINATION OF THE RELATIVE VOLUMES OF OXYGEN AND NITROGEN IN THE AIR (Optional)

Apparatus. Apparatus shown in Fig. 56. This is the same as shown in Fig. 43, Exercise 16, except that bottle *B* holds not less than 2 l. while bottle *C* holds about 1 l.; pneumatic trough; wing-top burner; cork to fit bottle *C*; graduated cylinder.

Materials. Copper gauze spiral used in Exercise 16. If the spiral is oxidized, reduce it as directed in 3, Exercise 16.

1. To determine the relative volume of oxygen and nitrogen in air, a definite volume of air is passed *slowly* over hot copper,

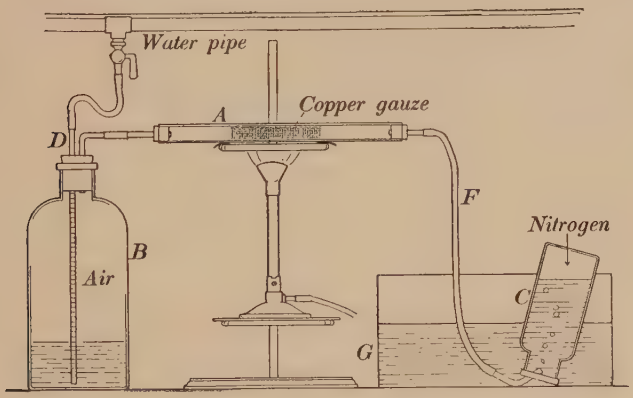

Fig. 56. Determining the relative volumes of oxygen and nitrogen in the air

which combines with the oxygen. The remaining nitrogen is collected and its volume measured. The experiment is the same as in Exercise 16 except that the volume of the air used and that of the resulting nitrogen are both measured.

Arrange the apparatus as shown in Fig. 56. The glass tube *A* contains the copper spiral. This spiral must be of such a size that it will just enter the tube and yet can be easily pushed to the middle part of the tube in the position shown in the figure. See that the corks and tubes are air-tight. At the beginning of the experiment bottle *B* is empty, while bottle *C* is *completely filled* with water and inverted in *G*. The end of tube *D* must touch the bottom of bottle *B*.

When the apparatus is all connected, heat the copper spiral in tube *A* with a wing-top burner until it is a dull-red heat, and proceed to prepare nitrogen exactly as in 1, Exercise 16. Care must be taken not to hasten the experiment, as some of the oxygen may not combine with the copper. When the bottle *C* is nearly filled, turn off the water, withdraw the heat, and disconnect the apparatus.

Now raise (or lower) the bottle *C* until the level of the water inside and outside the bottle is the same (it may be necessary to add more water to the trough); then, while holding it in this position, push the cork tightly into the mouth of the bottle. The bottle is then quickly raised from the trough and placed in an upright position on your desk.

By means of a graduated cylinder measure the volume of the water in *B*; this equals the volume of air analyzed. Record the volume in the table below. In the same way measure the volume of the nitrogen in *C* (this contains about 1 per cent of rare gases, chiefly argon) and record it in the table. Then from your results calculate the volumes of nitrogen and oxygen in 100 volumes of air, recalling that 100 volumes of air contains approximately 1 volume of the rare gases which remain mixed with the nitrogen.

Volume of air analyzed (volume of water in *B*) .	cc.
Volume of nitrogen and rare gases in the volume of air analyzed (volume of gas in *C*)	cc.
Volume of oxygen in air analyzed (difference between the two volumes above)	cc.

Volume of nitrogen and rare gases in 100 volumes of air (calculate)		cc.
Volume of nitrogen alone in 100 volumes of air (equals volume of nitrogen and rare gases in 100 volumes of air less 1 volume)		cc.
Volume of oxygen in 100 volumes of air (calculate)		cc.

	Results Obtained	Average of Results Obtained by Class
	Actual	
Volumes of oxygen in 100 volumes of air	21	
Volumes of nitrogen in 100 volumes of air	78	

2. (*a*) What are the sources of error in the above experiment? (*b*) Why is no attention paid to the temperature of the gases or to the barometric pressure?

EXERCISE 26

A STUDY OF SOLUTIONS

Apparatus. Test tubes; graduated tube or pipette; funnel and filter paper; 100-cc. beaker; watch glass; apparatus shown in Fig. 57 (*A* is a 250-cc. flask, *B* is a thermometer, and C is an open glass tube).

Materials. 2 crystals of potassium permanganate; 0.2 g. powdered calcium sulfate; 3 g. common salt; 3 g. potassium nitrate; 1 g. sugar; 10 cc. carbon tetrachloride (R. S.); 1 cc. oil (cottonseed or olive); 10 g. sodium thiosulfate; sufficient cotton to stopper a test tube.

1. Nearly fill two test tubes with water and set them in a rack. Drop into each a small crystal of potassium permanganate. Shake the contents of one tube and repeat the shaking after a few minutes, but do not move the other tube. (*a*) At the close of the laboratory period note the appearance of the liquid in each tube. (*b*) What do the results show?

2. Place exactly 3 g. of common salt in one test tube and an equal weight of potassium nitrate in another. Add to each exactly 1 cc. of water and heat each tube until the water boils. If the solid does not dissolve, add an additional cubic centimeter of water and again heat. Repeat until the solid in each

tube is dissolved, avoiding any excess of water. (*c*) Compare the approximate solubilities of the two solids in the boiling water.

(*d*) Cool the solutions in each tube and note the approximate amounts of solids separating. (*e*) Compare the results with the table of solubility of solids given in the Appendix of text.

3. Introduce about 0.5 g. of sugar into each of two test tubes. To the one tube add 5 cc. of water and to the other 5 cc. of carbon tetrachloride. (*f*) Shake the tubes gently and note the results. Repeat the experiment, substituting 3 or 4 drops of an oil for the sugar. (*g*) How could you remove from cloth a stain due to a sugar sirup (molasses)? to oils or grease? to a mixture of the two?

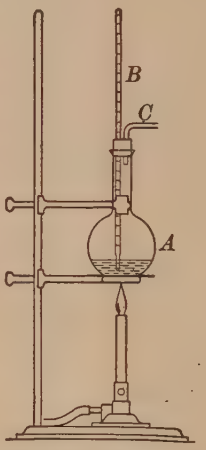

FIG. 57. Determining the boiling point of a salt solution

4. Introduce into a test tube 10 g. of sodium thiosulfate (ordinary "hypo" of the photographer) and add 2 cc. of water. Heat the tube gently until a uniform solution is obtained, care being taken that no particles of the solid remain on the side of the tube; stopper the tube loosely with a plug of cotton and set it aside until the solution is cold. If sufficient care has been taken, no solid will have separated. Now remove the cotton plug and drop into the solution a bit of the solid "hypo" as large as a pin point. (*h*) Note the results (hold the tube in a good light) and explain.

5. By means of the apparatus shown in Fig. 57, determine the boiling point of a saturated solution of common salt in water, keeping the bulb of the thermometer immersed in the solution. (*i*) Compare the boiling point of the solution with that of pure water.

EXERCISE 27

THE DETERMINATION OF THE SOLUBILITY OF COMMON SALT

Apparatus. 60-cc. bottle; evaporating-dish with watch-glass cover; beaker, tripod, and burner as shown in Fig. 58; funnel and filter paper; thermometer.

Materials. 15 g. common salt.

1. Introduce about 15 g. of finely powdered common salt into a 60-cc. bottle and add 40 cc. of water. Shake the mixture vigorously and set aside for about ten minutes, repeating the shaking several times at intervals of from one to two minutes so as to form a saturated solution. Take the temperature of the solution and record it in the table below.

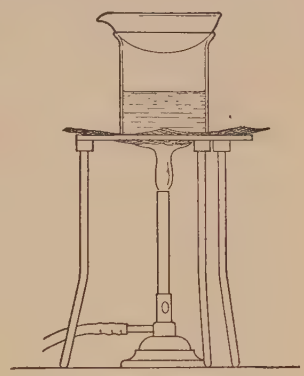

Fig. 58. Determining the solubility of common salt

Accurately weigh to 0.01 g. a small evaporating-dish and watch-glass cover and record the weights in the table; then filter into the dish about 20 cc. of the saturated solution of salt and reweigh, recording the weights in the table. Remove the watch glass and place the dish in a beaker partially filled with water as shown in Fig 58. Keep the water in the beaker boiling. After the solution in the dish has evaporated to dryness, remove the dish and place it on a ring stand (Fig. 24). Now cover the dish with the watch glass and heat it directly with the burner, regulating the flame so that the tip barely touches the dish. Continue the heating until all the moisture has been expelled and the under part of the watch glass is free from moisture. (*a*) Why use the watch-glass cover?

Now withdraw the burner, and after the dish is cool (room temperature) weigh the dish, residue, and cover, recording the weights in the table.

From the results so recorded, fill in the remaining blanks in the table and make the calculations called for. (*b*) Compare your results with those given in the Appendix of the text.

Temperature of the salt solution	°
Weight of empty dish and glass cover	g.
Weight of dish, salt solution, and glass cover	g.
Weight of salt solution taken	g.
Weight of dish, residue (salt) left after evaporation of water, and glass cover	g.
Weight of water present in solution (loss on evaporation)	g.
Weight of salt dissolved in the water (residue)	g.
Weight of salt that will dissolve in 100 g. of water at the temperature of the salt solution (calculate)	g.

EXERCISE 28

THE PREPARATION AND PROPERTIES OF CHLORINE

Apparatus. Test tube; apparatus as shown in Fig. 59 (*A* is a 250-cc. flask and *B* and *C* are 250-cc. bottles; *B* contains some sulfuric acid); three additional 250-cc. bottles (dry); glass plates.

Materials. 25 g. manganese dioxide; hydrochloric acid; sulfuric acid; bit of powdered antimony; strips of colored calico; piece of printed paper (printer's ink); paper written over with ordinary ink.

Precaution. *All the following experiments must be performed in the hood, and great care must be taken not to inhale the chlorine.*

1. Usual laboratory method of preparation. (Two students may work together.) Arrange an apparatus according to Fig. 59. Put into the flask from 20 to 25 g. of manganese dioxide. Insert the cork and pour 150 cc. of hydrochloric acid through the funnel tube. Shake the flask so as to mix the contents thoroughly. Warm gently, applying just enough heat to cause a gentle evolution of the gas, but not sufficient to boil

the liquid. Chlorine is set free and, escaping through F, bubbles through the sulfuric acid in B (which removes all moisture) and is collected in C. Fill three bottles with the gas (tell when filled by the yellowish color of the gas, which is best seen by placing a white cardboard back of the bottle). Cover the bottles with glass plates and set them aside.

Last of all, prepare some chlorine water by bringing the exit tube into a bottle containing 50 cc. of water so that the gas bubbles up through the liquid. Continue the gentle heating until no more chlorine is absorbed; then cork the bottle, label it, and set it aside in a dark place ((*a*) why dark?) for future use. (*b*) Write the equation for the reaction that takes place in the preparation of chlorine.

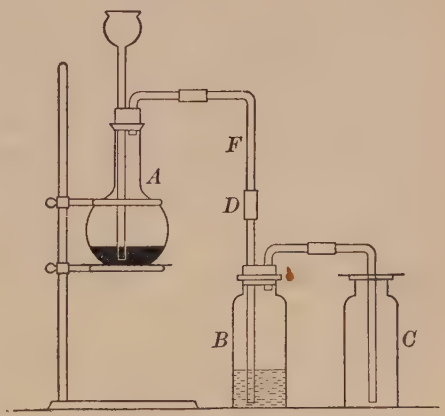

FIG. 59. The preparation of chlorine by the action of hydrochloric acid on manganese dioxide

2. Chemical conduct of chlorine. Grind a fragment of antimony to a fine powder and sprinkle a pinch of the powder into one of the bottles of the gas. $SbCl_3$ is formed. (*c*) Note the results and write the equation for the reaction.

Suspend strips of colored calico in a bottle of the gas; also two strips of paper, the one with writing in ink on it, the other with printing (printer's ink) on it. Dip similar strips of calico and paper into water until wet; then suspend these in another bottle of the gas.

(*d*) Record your results in each case. (*e*) What part does the water play in the bleaching?

3. (*f*) Summarize the properties and chemical conduct of chlorine.

EXERCISE 29

THE PREPARATION AND PROPERTIES OF HYDROGEN CHLORIDE AND OF HYDROCHLORIC ACID

Apparatus. Flask and bottle connected as shown in Fig. 60 (this is same as shown in Fig. 59, except that the bottle *B* contains water and the glass tube extending into the bottle *B does not touch the water in B*); two 250-cc. bottles (*dry*); large beaker; medicine dropper.

Materials. Dilute sulfuric acid prepared (*care*) by slowly pouring (1 or 2 cc. at a time with constant stirring) 18 cc. of the concentrated acid into 8 cc. water; 30 g. sodium chloride; splint; blue litmus paper.

1. Usual laboratory method of preparing hydrogen chloride. (Hood.) Arrange the apparatus as shown in Fig. 60. The bottle *B* contains water. *Care must be taken that the glass tube in the bottle B does not quite touch the surface of the water in the bottle.* Put about 30 g. of common salt into the flask *A*, insert the cork, pour the cold dilute sulfuric acid through the funnel tube, and mix the contents by a gentle motion of the flask; then connect the flask with *F* as shown in the figure. After two or three minutes warm gently with a *small* flame. Notice the currents in the water in *B*. (*a*) What causes them? When the gas is evolved regularly, disconnect the generator flask at *D* long enough to collect two bottles (dry) of the gas by displacement of air as in Fig. 53 (tell when filled by the fog formed at mouth of bottle). Cover these tightly with dry glass plates and set them aside; then connect the generator

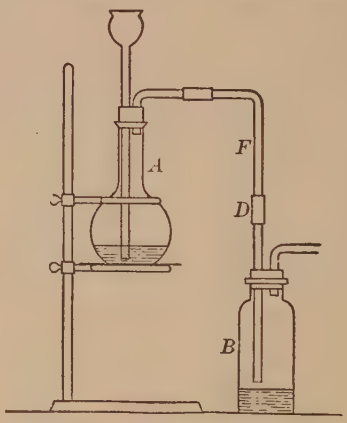

FIG. 60. The preparation of hydrogen chloride by the action of sulfuric acid on sodium chloride

with *B* again and continue to apply a gentle heat as long as any gas is evolved. (*b*) Write the equation for the reaction.

2. Properties and chemical conduct of hydrogen chloride. (*c*) What is the color of the gas (examine that in flask *A*)? Test the gas in one of the bottles with a lighted splint. (*d*) Is it combustible? (*e*) Is it a supporter of combustion? Fill a large beaker with water. Now uncover the remaining bottle, invert it, and at once bring its mouth under the surface of the water in the beaker and hold it in this position for two or three minutes. (*f*) Describe the results. (*g*) What does the experiment prove? (*h*) Why not extend the tube in bottle *B* (Fig. 60) to the bottom of the bottle?

3. Properties of hydrochloric acid. Put a drop of the solution from bottle *B* on a bit of blue litmus paper. (*i*) Note the result. (*j*) Pour 2 drops of the solution into 5 cc. of water and taste a drop. (*k*) Perform a test-tube experiment to prove the presence of chlorine in the acid; (*l*) also one to prove the presence of hydrogen. (*m*) How does the solution compare with the hydrochloric acid on your desk?

4. (*n*) State clearly the difference between the terms *hydrogen chloride* and *hydrochloric acid*.

EXERCISE 30

SODIUM; SODIUM HYDROXIDE

Apparatus. Evaporating-dish with glass-plate cover; forceps; glass rod.

Materials. Bit of sodium as large as a pea; red litmus paper.

1. Properties of sodium. Recall experiment 1, Exercise 9. Obtain from your instructor a bit of sodium. Hold the sodium with the forceps on a glass plate, cut it, and note the rapidity with which the freshly cut surface is tarnished. Half fill your evaporating-dish with water; then drop into this a bit of sodium as large as a pea, quickly cover the dish with the glass plate, and leave it covered until the sodium has all disappeared.

(*a*) Is sodium heavier or lighter than water? (*b*) Write the equation for the reaction between sodium and water. (*c*) What is the composition of the liquid in the evaporating-dish?

2. Properties of sodium hydroxide. Place a drop of the liquid in the evaporating-dish on a piece of red litmus paper. (*d*) Contrast its action with that of hydrochloric acid on litmus paper (Exercise 29). Mix 1 or 2 drops of the solution with 5 cc. of water and taste a drop of the resulting solution. (*e*) Contrast with the taste of hydrochloric acid (Exercise 29). Evaporate the solution to dryness. (*f*) What is the residue?

EXERCISE 31

THE PROPERTIES OF ACIDS, BASES, AND SALTS

Apparatus. Small beaker; stirring-rod; evaporating-dish; ring stand; test tubes; medicine dropper.

Materials. A few drops of each of the following acids: hydrochloric, sulfuric, nitric, acetic; solutions of the following bases: sodium hydroxide, potassium hydroxide (R. S.), calcium hydroxide (R. S.); strips of blue and of red litmus paper; hydrochloric acid.

1. Acids. Recall the properties of hydrochloric acid (Exercise 29). Prepare a dilute solution of each of the following acids by adding 2 or 3 drops of the acid to about 10 cc. of water in a test tube and mixing thoroughly: hydrochloric, sulfuric, nitric, acetic.

By means of a clean glass rod transfer a drop of each of the dilute solutions to a piece of blue litmus paper. (*a*) Note the result in each case. (*b*) Taste *one* drop of the dilute solutions (*rinse the mouth with water after tasting*).

Compare the formulas of the acids. (*c*) In what respect are the acids similar in composition?

2. Bases. In a similar way try the effect on red litmus paper of a solution of each of the following bases: sodium hydroxide (recall Exercise 30), potassium hydroxide, calcium hydroxide. (*d*) Do they affect the blue litmus paper? (*e*) Taste a drop of the calcium hydroxide solution.

Compare the formulas of the bases. (*f*) In what respect are the bases similar in composition?

3. Salts. Dilute 5 cc. of the ordinary laboratory solution of sodium hydroxide (1 part of the hydroxide by weight to 10 parts of water) with an equal volume of water. To this solution add 4 or 5 drops of the ordinary concentrated hydrochloric acid. Stir the resulting solution with a glass rod and test its action on blue and on red litmus paper. (*g*) Has it acid or basic properties?

Now continue to add the acid drop by drop until the resulting solution is neutral (that is, has no effect on either blue or red litmus paper) or is, at most, slightly acid. (*h*) Write the equation for the reaction between the hydroxide and the acid. Now pour the solution into an evaporating-dish and evaporate to dryness. (*i*) What compound remains? (*j*) Taste it.

(*k*) What is the name given to the compounds formed by the action of acids on bases?

4. (*l*) Characterize acids and bases as to composition; as to their action on litmus; as to taste; as to their interaction with each other.

EXERCISE 32

THE RATIO OF ACID TO BASE IN NEUTRALIZATION

Apparatus. 2 burettes and supports, as shown in Fig. 61; 2 small beakers and stirring-rod; graduated pipette.

Materials. Sodium hydroxide solution prepared by adding 20 cc. of the laboratory reagent to 100 cc. water; 1 cc. sulfuric acid (use graduated pipette) added to 100 cc. water and mixed thoroughly; a few drops of a phenolphthalein solution (R. S.).

(Two students may work together.)

1. Rinse out a burette, first with distilled water and then with a little of the solution of sodium hydroxide. Support the burette (Fig. 61) and pour into it the hydroxide solution until the level of the liquid is 1 or 2 cm. above the zero mark. Turn the stopcock and let the solution slowly flow out until the bot-

tom of the curved surface (Fig. 3) of the liquid in the burette is on a level with the zero mark. In a similar way fill a second burette with the acid solution.

Now let exactly 15 cc. of the acid solution flow into a small beaker, add two drops of phenolphthalein solution, and run in 2 or 3 cc. of the hydroxide solution. Notice that where the liquids come in contact a reddish color is produced, which disappears quickly on stirring. Run in more of the solution, a little at a time, until the color fades slowly, and then a drop at a time until the entire liquid, on stirring, remains colored *faintly* pink. This marks approximately the point of neutralization. Record the number of cubic centimeters of the hydroxide solution used in the table below. Repeat the experiment, using 30 cc. of the acid, and again record in the table the number of cubic centimeters of the hydroxide solution used to neutralize the acid.

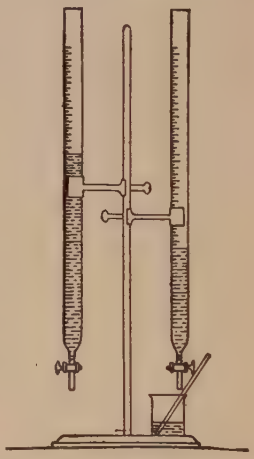

FIG. 61. Graduated tubes (burettes) for measuring volumes of liquids

First Experiment

Volume of hydroxide solution used to neutralize 15 cc. of the acid . . . . . cc.

Volume of hydroxide solution required to neutralize 1 cc. of acid (calculated) . . . . . cc.

Second Experiment

Volume of hydroxide solution used to neutralize 30 cc. of the acid . . . . . cc.

Volume of hydroxide solution required to neutralize 1 cc. of acid (calculated) . . . . . cc.

2. (*a*) What do the results of the experiments prove?

EXERCISE 33

CARBONIC ACID AND ITS SALTS (CARBONATES)

Apparatus. Hydrogen generator, as used for preparing carbon dioxide in Exercise 23; small beaker; 5 test tubes.

Materials. Pieces of marble for generating carbon dioxide; hydrochloric acid; blue litmus paper; 5 cc. sodium hydroxide solution diluted with 10 cc. water; 1 g. of each of the common carbonates, such as sodium carbonate, magnesium carbonate, calcium carbonate; 25 cc. limewater (R. S.).

1. When carbon dioxide is passed into water, a portion of the gas combines with the water to form a compound known as carbonic acid ($H_2O + CO_2 \longrightarrow H_2CO_3$). This is a very weak acid and so unstable that it can be obtained only in dilute solutions. If an attempt is made to concentrate the solution by evaporating the water, the acid decomposes again into water and carbon dioxide. The salts of this acid (the *carbonates*), on the other hand, are stable, and many of them are important and common compounds. Thus, calcium carbonate ($CaCO_3$) is the chief constituent of limestone, while sodium carbonate (Na_2CO_3) is common washing soda.

Generate carbon dioxide (Fig. 53) and pass the gas through 25 cc. of water. (*a*) Taste the liquid. (*b*) Is the acid formed strong enough to affect blue litmus paper?

2. Pass carbon dioxide through 5 cc. of a solution of sodium hydroxide until the gas is no longer absorbed (the carbon dioxide combines with the water present to form carbonic acid. This acid then reacts with the base, sodium hydroxide, to form a salt and water). Evaporate the solution to dryness. (*c*) What is the product left in the dish? (*d*) Write the equation for its formation. (*e*) Could a solution of sodium hydroxide be used in place of a solution of calcium hydroxide (limewater) in testing for carbon dioxide?

3. Examine the physical properties of such carbonates as are available. (*f*) What ones are soluble in water (p. 304 of text)? Test the action of hydrochloric acid or sulfuric acid on each by adding 1 or 2 drops of the acid to 0.1 g. of the carbonate in a test tube. What evidences have you that a gas is evolved? (*g*) Test the gas to determine whether or not it is carbon dioxide (Fig. 54). *All carbonates evolve carbon dioxide when treated with hydrochloric or sulfuric acid. This reaction serves as a good test for carbonates.*

EXERCISE 34

A METHOD FOR DETERMINING WHETHER A GIVEN LIQUID IS A CONDUCTOR OF ELECTRICITY (Optional)

Apparatus. Current from electric-lighting system; apparatus as shown in Fig. 62. *B* is a plug for connecting the apparatus with any ordinary electric-lighting system. *C* is an ordinary incandescent lamp. *A* is a small bottle (60 cc.). The wires are the common insulated copper wires, but the insulation is removed from that portion of the wires which extend inside the bottle *A*. The apparatus can be purchased from supply houses, but is easily made.

Materials. 5 g. common salt; 5 g. sugar; 5 cc. sodium hydroxide solution added to 20 cc. water; 3 cc. sulfuric acid added to 20 cc. water; tap or well water; 10 cc. hydrochloric acid; (10 cc. of a benzene solution of hydrogen chloride).

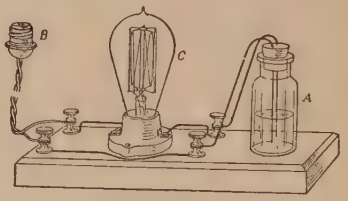

Fig. 62. Apparatus for determining whether or not a liquid is a conductor of electricity

1. Obtain from your instructor the apparatus shown in Fig. 62. Polish the ends (electrodes) of the copper wires that extend into the bottle *A* with emery paper until they are bright and free from oxide. *At the beginning of each experiment see that the electrodes are bright and dry and that the bottle is also perfectly clean and dry.* Unscrew a lamp *C* from a convenient socket in the laboratory, screw it *loosely*

into the socket on your apparatus, and attach the apparatus to the empty socket on the lighting system by means of the extension cord and plug *B*. Every time a change is to be made in the cell, loosen the lamp *C* in the socket, and do not screw it down to make contact until all the connections of the cell have been arranged.

2. Partly fill the bottle *A* with dry, powdered salt, dip the electrodes into the powder, arrange the connections at the binding-posts, and screw down the lamp *C*. (*a*) Does the salt conduct the electric current?

3. (*b*) Repeat 2, using distilled water.

4. (*c*) Repeat 2, using a solution of the salt in distilled water.

5. (*d*) Test the conductivity of the following substances and interpret the results: dry, powdered sugar; a solution of sugar in water; tap or well water; distilled water containing a few drops of sulfuric acid; distilled water containing a few drops of hydrochloric acid; a solution of sodium hydroxide. (*e*) How do we account for the fact that solutions of some compounds conduct the electric current while others do not?

6. Secure from your instructor 10 cc. of a benzene solution (*benzene is inflammable*) of hydrogen chloride (easily prepared just as an aqueous solution is prepared, as described in Exercise 29). (*f*) Test its conductivity. (*g*) Test its effect on blue and on red litmus paper. (*h*) Does it dissolve zinc? (*i*) How do you account for the difference in properties between a benzene solution of hydrogen chloride and an aqueous solution of the same gas?

7. (*j*) Define the terms *acid, base,* and *salt* from the standpoint of the ionization theory.

EXERCISE 35

THE DISPLACEMENT OF METALS FROM THEIR COMPOUNDS (THE DISPLACEMENT SERIES)

Apparatus. 4 test tubes; test-tube rack.

Materials. 4 clean and bright strips each of zinc and copper (1 cm. × 10 cm.); 0.5 g. lead nitrate dissolved in 10 cc. water; 0.5 g. copper nitrate dissolved in 10 cc. water; 0.5 g. mercuric nitrate dissolved in 10 cc. water; 3 cc. sulfuric acid dissolved in 10 cc. water.

1. Pour into separate test tubes to a depth of 4 or 5 cm. solutions of the following compounds: lead nitrate, dilute sulfuric acid, copper nitrate, mercuric nitrate. Set the tubes in a rack in the order given above and label them *A*, *B*, *C*, and *D respectively* (Fig. 63).

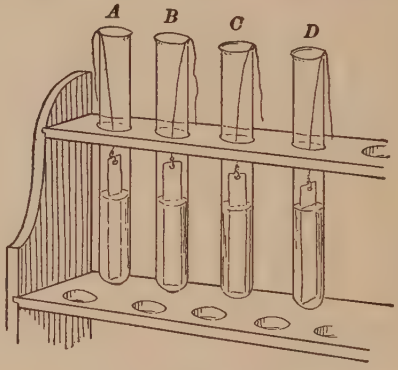

FIG. 63. Testing the action of different metals on solutions of salts

Now place in each tube a strip of zinc. (It is convenient to have a strong thread attached to the upper part of each strip so that the strip may easily be withdrawn from the tube. The strips should be only partly immersed in the solution.) Note any change taking place in the appearance of the zinc.

After twenty minutes withdraw the strips and wipe them on a piece of white paper. (*a*) Note any evidence tending to show that the zinc has displaced the lead, hydrogen, copper, and mercury from their compounds. (Metals in a very finely divided form are black, as a rule.) (*b*) Account for any change in the color of the copper nitrate solution.

2. Repeat experiment 1, substituting for the zinc a strip of copper. (*c*) Contrast the results obtained with those obtained in 1. (*d*) How do you account for the change in the color of the solution of mercuric nitrate after the addition of the copper strip? (*e*) Are your results in accord with the table of the displacement series given on page 161 of text?

EXERCISE 36

THE PREPARATION AND PROPERTIES OF AMMONIA

Apparatus. Test tube; hard-glass test tube fitted with cork and tubing as shown in Fig. 64; graduated tube; three 250-cc. wide-mouthed bottles; 2 pieces of window glass; large beaker; glass rod.

Materials. Solution of sodium hydroxide (the ordinary desk solution will serve); 9 g. ammonium chloride; 15 g. powdered calcium hydroxide (slaked or hydrated lime, $Ca(OH)_2$); strip of red litmus paper; strip of blue litmus paper; hydrochloric acid; splint.

1. Dissolve 0.5 g. of ammonium chloride in 3 or 4 cc. of water in a test tube and heat to boiling. (*a*) Note the odor.

Now add 3 cc. of a solution of sodium hydroxide to the hot solution of ammonium chloride and continue the heating. (*b*) Again note the odor. (*c*) Moisten a strip of red litmus paper and hold it at the mouth of the tube but not in contact with it.

(*d*) Dip the end of a glass rod in a concentrated solution of hydrochloric acid and hold it in the mouth of the test tube. Dense white fumes of ammonium chloride (NH_4Cl) are formed.

(*e*) Complete the following equations:

$$NH_4Cl + NaOH \longrightarrow$$
$$NH_3 + H_2O \longrightarrow$$
$$NH_4OH + HCl \longrightarrow$$

2. **Usual laboratory method for preparing ammonia.** This differs from the method used in 1 only in the fact that the less expensive calcium hydroxide (slaked lime) is substituted

for the sodium hydroxide. The form of apparatus used is shown in Fig. 64. The bottle *B* (250-cc.) contains 25 cc. of water. The glass tube *C* extends through a hole in a cardboard resting on the mouth of the bottle. The end of the tube must *just touch* the water in the bottle.

Place in the tube *A* a mixture of 15 g. of powdered slaked lime and 8 g. of ammonium chloride. Connect the tube as shown in Fig. 64 and heat the mixture gently, beginning with that portion near the mouth of the tube and gradually extending the heat to the other portions. As soon as the gas is evolved freely (as shown by the bubbles at the end of the tube *C*), bring the tube *C* to an upright position, as shown in the dotted lines, and collect two bottles of the gas. To do this, bring the bottles successively down over the exit tube, leave each in this position until a drop of hydrochloric acid on the end of a glass rod fumes strongly when held at the mouth of the bottle; then withdraw the bottle, cover its mouth with a glass plate, and set it aside, mouth downward.

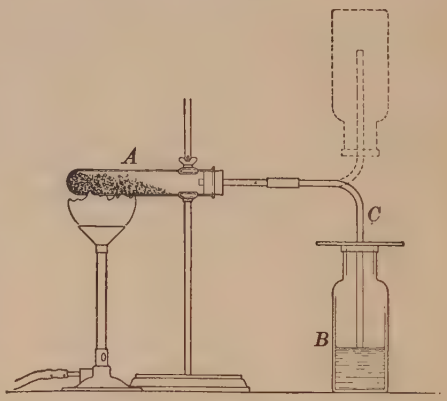

FIG. 64. Preparing ammonia by heating a mixture of ammonium chloride and calcium hydroxide (slaked lime)

When both bottles are filled, bring the tube *C* into the bottle *B* again and continue to heat the mixture gently as long as any gas is generated. (*f*) Write the equations for all the reactions involved.

3. Properties and chemical conduct of ammonia. (*g*) Note the color and odor of the gas. (*h*) Is it heavier or lighter than air? (*i*) Test a bottle of the gas with a burning splint.

Devise a simple experiment for finding out whether or not ammonia is soluble in water. After your method is approved by the instructor, try it out with the remaining bottle of the gas. (*j*) Describe the results.

4. Properties of ammonium hydroxide. (*k*) Note the odor of the liquid in the bottle *B*. (*l*) Try its effect on blue and on red litmus paper. (*m*) How does it compare with the aqua ammonia of the druggist in its odor and its action on litmus? (*n*) Does the gas combine with the water or is it simply dissolved in the water? (*o*) Give reasons for your answer. Now neutralize the liquid with hydrochloric acid and evaporate just to dryness (Fig. 58). (*p*) Compare the residue with the ammonium chloride used in experiment 1.

5. (*q*) Summarize the properties and chemical conduct of ammonia.

EXERCISE 37

THE PREPARATION AND PROPERTIES OF NITRIC ACID

Apparatus. Glass retort (150-cc.), test tube, and beaker (500-cc.), arranged as shown in Fig. 65; funnel; evaporating-dish.

Materials. 12 g. sodium nitrate; 10 cc. sulfuric acid; small piece of tin; small strip of copper.

Caution. *The students must remember that both sulfuric acid and nitric acid are very corrosive and must exercise care in handling them.*

1. Preparation of nitric acid. Arrange an apparatus like that shown in Fig. 65. Put in the retort *A* about 12 g. of sodium nitrate and 10 cc. of sulfuric acid, pouring the latter through a funnel placed in the tubulus *B* of the retort. Heat the mixture gently with a small flame. Nitric acid is set free, distills over, and is condensed in the test tube *C*, which is kept cold by being partly immersed in ice water in the beaker *D*. (*a*) Write the equation for the reaction between the sulfuric acid and the sodium nitrate.

2. Properties and chemical conduct of nitric acid. When nitric acid is heated, a part of it is decomposed into water, nitrogen dioxide, and oxygen. On this account it is a good oxidizing agent. To test its oxidizing properties, put a small piece of tin in a test tube, cover it with a little nitric acid, and gently heat (hood). The white product formed is composed mainly of tin and oxygen, the latter being supplied by the nitric acid.

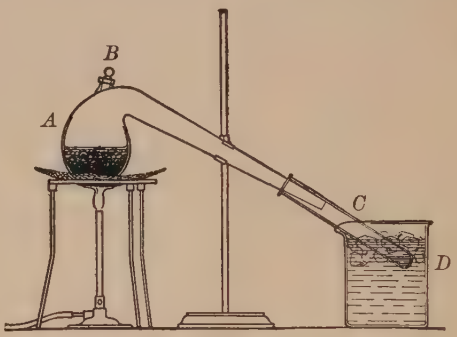

FIG. 65. The preparation of nitric acid by the action of sulfuric acid on sodium nitrate

Pure nitric acid is colorless. (*b*) How do you account for the color of the acid which you have prepared?

Place a small strip of copper in an evaporating-dish (hood) and add some of the acid you have prepared, a few drops at a time, until the copper is just dissolved. Evaporate the solution to dryness (Fig. 58). (*c*) Note the appearance of the residue. (*d*) Since copper is below hydrogen in the displacement series, how do you account for the fact that nitric acid dissolves the metal (consult text)? Save the residue in the dish ((*e*) what is it?) for use in the following exercise.

EXERCISE 38

THE PROPERTIES OF THE SALTS OF NITRIC ACID (NITRATES)

Apparatus. Evaporating-dish containing the copper nitrate prepared in Exercise 37; 5 test tubes in test-tube rack; medicine dropper.

Materials. Strip of copper; crystal of lead nitrate; crystals of such nitrates as are available, including sodium nitrate and potassium nitrate; sulfuric acid; 2 g. ferrous sulfate dissolved in 10 cc. water.

1. Properties of nitrates. Heat the dish (hood) containing the copper nitrate, prepared in Exercise 37, with a small flame. (*a*) Note the color of the gas evolved, also the color of the residue. (*b*) Compare the residue with the crust of copper oxide prepared by holding a strip of copper in the tip of a flame.

Place a crystal of lead nitrate in the evaporating-dish and heat gently. (*c*) Compare with the results obtained in 1.

Place a small crystal of such nitrates as are available in your laboratory in separate test tubes and test their solubility in water. (*d*) What nitrates are insoluble in water (consult text)?

2. How to detect the presence of a nitrate. Dissolve a crystal of sodium nitrate in 2 or 3 cc. of water in a test tube, add drop by drop an equal volume of sulfuric acid, mixing the liquids as the acid is added, then cool the mixture. The sulfuric acid acts on the nitrate, liberating nitric acid. Now tip the tube slightly and gently pour 2 or 3 cc. of the solution of ferrous sulfate down the side of the tube, so that it floats on the heavier liquid, and set the tube aside, being careful not to mix the two liquids. A brown ring soon forms where the liquids meet. Repeat the experiment, using potassium nitrate. This is a good test for nitrates. The brown ring is due to the presence of a compound formed by the action of ferrous sulfate on nitric acid.

Secure one or more compounds from your instructor and test to see if they are nitrates. (*e*) Record your results.

EXERCISE 39

THE PREPARATION AND PROPERTIES OF SOME OF THE OXIDES OF NITROGEN

Apparatus. Hard-glass test tube, with delivery tube, as used in preparing oxygen (Fig. 28); 3 wide-mouthed bottles (250-cc.); pneumatic trough; hydrogen generator (Fig. 35); 2 glass plates.

Materials. 8 g. ammonium nitrate; wooden splints; 5 small strips of copper; 10 cc. nitric acid.

1. Nitrous oxide. Put 6 or 8 g. of ammonium nitrate in the hard-glass test tube used in the preparation of oxygen (Fig. 28). Attach a delivery tube and heat *gently*, applying no more heat than is necessary to cause a slow evolution of the gas.

As soon as the gas is regularly evolved, collect two or three bottles of it over water. Notice the water deposited on the sides of the test tube. (*a*) What is the source of it? (*b*) Note the color, odor, and taste of the gas. (*c*) Test it with a glowing splint. (*d*) Account for the result. (*e*) How can you distinguish between nitrous oxide and oxygen?

2. Nitric oxide and nitrogen dioxide. Put a few pieces of copper in your hydrogen generator (hood) (Fig. 35), just cover them with water, and add 2 or 3 cc. of nitric acid. Collect over water two bottles of the evolved gas, adding more nitric acid to the liquid in the generator if necessary to secure the required amount of gas.

(*f*) Compare the color of the gas in the generator with that collected in the bottles and account for any difference. (*g*) Write the equations for all the reactions involved.

(*h*) Uncover one of the bottles of the gas and account for the result. (*i*) Test the gas in the second bottle with a burning splint. (*j*) Which is the more stable, nitrous oxide or nitric oxide? (*k*) Give reasons for your answer.

EXERCISE 40

THE PROPERTIES AND CHEMICAL CONDUCT OF SULFUR

Apparatus. 3 test tubes; smallest-sized beaker; magnifying-glass; porcelain crucible; pipestem triangle; large beaker; steel forceps.

Materials. 5 cc. carbon disulfide; 20 g. powdered brimstone; strip of copper foil (0.5 × 3 cm.); 5 g. iron powder.

1. Properties of sulfur. (*a*) Examine the physical properties of a piece of brimstone. Pour 2 or 3 cc. of carbon disulfide (hood) over 2 g. of powdered brimstone in a test tube. (*Keep carbon disulfide away from flame and do not inhale the vapor.*) Cover the mouth of the tube with the thumb and shake the contents gently until the sulfur is dissolved, adding more carbon disulfide if necessary. Pour the clear solution into a small beaker, cover the beaker loosely with a filter paper, and set it aside in the hood. The carbon disulfide soon evaporates, the sulfur being deposited in crystals. (*b*) Examine these with a magnifying-glass, noting their general shape.

Half fill a test tube with powdered brimstone and heat it gently until the sulfur is just melted. (*c*) Note the properties of the liquid. Now apply a stronger heat and observe that the liquid becomes darker, and at a certain temperature (200°–250°) is so thick that the tube may be inverted without spilling it. Finally, increase the heat until the sulfur boils (448°), and then slowly pour the boiling liquid into a beaker of cold water. (*d*) Examine the product. (*e*) What name is given to this form of sulfur? Expose it to the air for an hour or longer. (*f*) Have its properties remained unchanged?

Fill a porcelain crucible with powdered brimstone and apply a very gentle heat until the sulfur is just melted. Withdraw the flame and examine the liquid carefully as it cools. Crystals soon begin to form on the surface of the melted sulfur, rapidly extending from the circumference toward the center. Before

they reach the center grasp the crucible with the forceps and quickly pour off the remaining liquid and examine the crystals. (*g*) Compare them in shape with those deposited from the solution of carbon disulfide. (*h*) In how many forms have you obtained sulfur?

2. **Chemical conduct of sulfur.** Burn a small piece of sulfur. (*i*) Notice the appearance of the burning sulfur and the odor of the gas (SO_2) formed.

Boil a little sulfur in a test tube and drop a small strip of hot copper foil (Fig. 66) in the boiling liquid. (*j*) Is there any visible evidence of a chemical change? (*k*) What is formed?

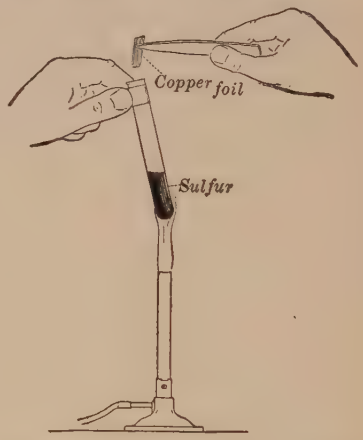

FIG. 66. Dropping a piece of copper foil into boiling sulfur

Grind together 5 or 6 g. of iron powder with an equal weight of sulfur, transfer the mixture to a test tube, and heat it strongly in a Bunsen flame. (*l*) Describe the results. Retain the solid formed for use in the following exercise.

3. (*m*) Summarize the properties and conduct of sulfur.

EXERCISE 41

THE PREPARATION AND PROPERTIES OF HYDROGEN SULFIDE

Apparatus. Hydrogen generator and tubes, as shown in Fig. 67; 2 wide-mouthed bottles (250-cc. and 60-cc.); funnel; evaporating-dish.

Materials. 10 g. ferrous sulfide (in pieces as big as a bean or larger); 20 cc. hydrochloric acid added to 20 cc. water; blue and red litmus papers; silver coin; filter paper.

1. (Hood.) Attach a delivery tube to the hydrogen generator, as shown in Fig. 67. Put into the generator *A* a few

pieces of ferrous sulfide (FeS) and insert the stopper. Now pour a little water through the funnel tube of the generator until the end of the tube just dips below the surface of the water; then pour in a few cc. of the hydrochloric acid, adding more from time to time, if necessary, to maintain a gentle evolution of the gas. (*a*) Write the equation. The gas escapes into the bottle *B*, which gradually becomes filled. (*b*) Very cautiously note the odor (***Caution.*** *The gas is poisonous if inhaled in concentrated form*); (*c*) also note the color of the evolved gas. Continue the evolution of the gas until it is ignited by a flame held at the mouth of the bottle *B*. (*d*) Account for the deposit formed on the sides of the bottle when the gas burns.

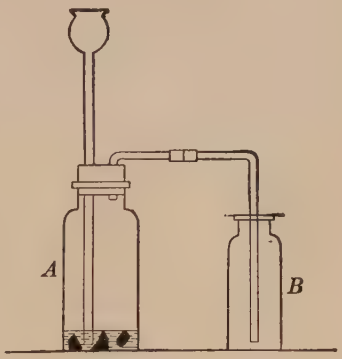

FIG. 67. Preparing hydrogen sulfide by the action of hydrochloric acid on ferrous sulfide

2. Replace the bottle *B* with a 60-cc. bottle half filled with water, and allow the gas from the generator (add more acid if necessary) to bubble through the water for one or two minutes. Test the resulting solution with blue and with red litmus paper. (*e*) What is the solution called? (*f*) How does it compare with the so-called "sulfur water" of many springs? (*g*) Drop a silver coin into the solution and account for the results. (*h*) Why do certain foods blacken silver spoons?

3. Transfer to a test tube a piece of the solid formed in 2, Exercise 40, by heating a mixture of iron and sulfur and add 1 or 2 cc. of dilute hydrochloric acid. (*i*) What gas is evolved (cautiously note the odor)? (*j*) Account for its formation.

4. (*k*) Summarize the properties of hydrogen sulfide.

EXERCISE 42

THE PREPARATION AND PROPERTIES OF THE SALTS OF HYDROSULFURIC ACID (SULFIDES)

Apparatus. Hydrogen generator and connections, as shown in Fig. 68; 6 test tubes; beaker; funnel; watch glass; iron spoon.

Materials. Ferrous sulfide and dilute hydrochloric acid, as used in Exercise 41; separate solutions of silver nitrate, copper, sulfate, cadmium chloride, lead nitrate, and sodium chloride, made by dissolving about 0.5 g. of the solid in 5 cc. water (solutions on reagent shelf may be used); sulfuric acid; lead acetate (R. S.); 5 g. sulfur; 3 g. lime.

1. Preparation of sulfides. (Hood.) Charge the hydrogen sulfide generator as in Exercise 41 and pass a few bubbles of the gas (Fig. 68) through each of the following solutions: silver nitrate, copper sulfate, cadmium chloride, sodium chloride, lead nitrate. The exit tube *C*, through which the gas bubbles into the solutions, must be thoroughly cleaned each time. (*a*) Note the color of the precipitate obtained in each case. Write the equations for the reactions involved. (*b*) Do any of the solutions fail to give a precipitate? (*c*) How do you account for this?

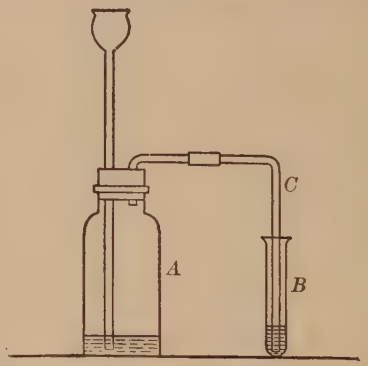

FIG. 68. Preparing sulfides of the metals by passing hydrogen sulfide through a solution of their salts

Intimately mix 5 g. of sulfur with 3 g. of powdered lime. Transfer to a beaker and add 150 cc. of water. Stir the mixture and heat just to boiling for ten minutes. Now fill a test tube with the resulting mixture, cork the tube loosely, and set it aside until the beginning of the next laboratory period; then examine. (*d*) Describe the

results. (*e*) For what is the solution used (consult text)? (*f*) What is its composition (consult text)?

2. **Test for hydrogen sulfide.** Dip a strip of filter paper into a solution of lead acetate. Remove the cork from the hydrogen sulfide generator and insert the paper for a moment. (*g*) Note and account for the results. This serves as a convenient test for the gas. (*h*) What property would also serve to detect the gas if present in any marked quantity?

3. **Test for sulfides.** Prepare some dilute sulfuric acid by adding (*care*) 2 or 3 drops of the acid on your desk to 1 cc. of water. Place a small bit of ferrous sulfide (FeS) on your watch glass and moisten it with the dilute acid. Cautiously note the odor. Most of the sulfides when treated in this way evolve hydrogen sulfide, which can be detected by its odor and by its action on paper moistened with lead acetate solution.

All sulfides when heated in air evolve sulfur dioxide (formed by the combustion of the sulfur present), which has the characteristic odor of burning sulfur. (*i*) Heat a little ferrous sulfide on an iron spoon in the flame of the burner and note the odor.

EXERCISE 43

SULFUR DIOXIDE AND SULFUROUS ACID

Apparatus. 250-cc. flask fitted with funnel tube and glass exit tube, as shown in Fig. 69; 3 bottles (250-cc.); watch glass; medicine dropper.

Materials. 10 g. copper; sulfuric acid; splint; blue litmus paper; crystal of sodium sulfite; strips of colored calico or a red flower.

1. (*a*) By what method have you already prepared sulfur dioxide?

2. **Preparation of sulfur dioxide and sulfurous acid by the action of sulfuric acid on copper (laboratory method).** Place about 10 g. of copper turnings or small pieces of sheet copper in a generator arranged as in Fig. 69. Add 25 cc. of concen-

trated sulfuric acid and apply a gentle heat. As soon as the action begins, lower the flame, regulating it so as to obtain a uniform evolution of the gas. Collect two bottles of the gas by displacement of air; then cause it to bubble through 25 cc. of water as long as any is dissolved. (*b*) Write the equation for the reaction of sulfuric acid on copper.

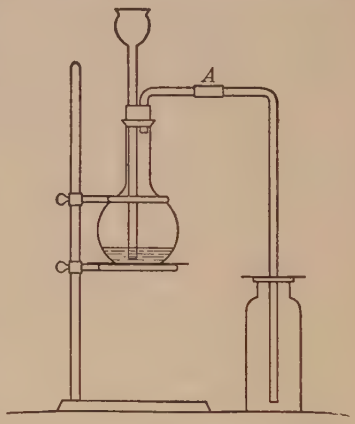

FIG. 69. Preparing sulfur dioxide by the action of sulfuric acid on copper

3. Properties and chemical conduct of sulfur dioxide and sulfurous acid. (*c*) Note the odor of the gas. (*d*) Is the gas combustible? (*e*) Test with blue litmus paper the liquid formed by passing the gas through water. (*f*) Does the gas combine with the water or simply dissolve in it? (*g*) Give reasons for your answer. (*h*) Immerse in the liquid some small strips of colored calico or some petals of a red flower and note any results.

4. Salts of sulfurous acid; the sulfites. Place a small crystal of some sulfite, such as sodium sulfite (Na_2SO_3), on a watch glass and moisten it with 2 or 3 drops of sulfuric acid. (*i*) Note the odor of the evolved gas. *All sulfites evolve sulfur dioxide when treated with sulfuric acid.* This reaction serves as a good test for sulfites. It also serves as a method for preparing sulfur dioxide.

5. (*j*) Enumerate three methods for preparing sulfur dioxide.

6. (*k*) Briefly summarize the properties and chemical conduct of sulfur dioxide.

EXERCISE 44

A STUDY OF SULFURIC ACID

Apparatus. Stirring-rod; test tubes; medicine dropper.

Materials. Sulfuric acid; 0.5 g. sugar; wooden splint; a small piece of zinc; piece of charcoal the size of a pea; barium chloride solution (R.S.); hydrochloric acid.

1. Properties and chemical conduct of sulfuric acid. By means of a glass rod place 1 or 2 drops of sulfuric acid on a wooden splint, and after two or three minutes (*a*) note the results produced. Put about 0.5 g. of sugar in a test tube and add 3 or 4 drops of the acid. (*b*) Note and account for the change produced.

Heat a bit of charcoal in a test tube with 1 or 2 cc. of concentrated sulfuric acid. (*c*) What gas is evolved (odor)? (*d*) Account for its formation, recalling that carbon has a strong affinity for oxygen.

In the preparation of hydrogen (Exercise 9) the directions called for dilute sulfuric acid. (*e*) Determine whether the concentrated acid will do as well, by pouring 2 or 3 cc. of the acid over a bit of zinc in a test tube (if no action takes place, heat the mixture gently and test the gas evolved both by a lighted splint and by noting its odor). (*f*) Account for the difference in the action on zinc between the dilute and the concentrated acid.

Recall the action of sulfuric acid on sodium chloride (Exercise 29) and on sodium nitrate (Exercise 37). (*g*) What property of sulfuric acid enables it to be used in the preparation of other acids?

2. Test for sulfuric acid. Add 3 drops of sulfuric acid (*care*) to 5 cc. of water in a test tube. To this add a few drops of a solution of barium chloride. The chloride reacts with the acid to form barium sulfate ($BaSO_4$), which is insoluble and hence

separates, as fast as formed, as a white precipitate. (*h*) Write the equation for the formation of the barium sulfate. Now add to the precipitate 4 or 5 drops of hydrochloric acid. (*i*) Does the precipitate dissolve? *The formation of a white precipitate with barium chloride, insoluble in hydrochloric acid,* constitutes a good test for sulfuric acid.

EXERCISE 45

SALTS OF SULFURIC ACID (SULFATES)

Apparatus. 6 test tubes.

Materials. Crystals or small amounts (0.1 g.) of the sulfates available in the laboratory; 2 cc. barium chloride solution (R.S.); hydrochloric acid.

1. Examine the physical properties of such sulfates as are available. Test the solubility of each in water by adding about 0.5 g. of the sulfate to a test tube nearly filled with water and shaking the mixture. (*a*) What sulfates are insoluble (consult text)?

2. Prepare a dilute solution of different sulfates by dissolving a crystal of each in 2 or 3 cc. of water. Add to each 1 drop of barium chloride solution. Now add 2 drops of hydrochloric acid to the mixture in each tube. (*b*) Does the precipitate dissolve? *All soluble sulfates give in solution a white precipitate ($BaSO_4$) with barium chloride solution, which precipitate is insoluble in hydrochloric acid.* This reaction serves as a good test for sulfates.

It will be noted that both sulfuric acid and its salts give with barium chloride the same product; namely, a white precipitate of barium sulfate. This is evident from the following facts: It will be recalled that both acids and salts are ionized in solution. In the case of sulfuric acid and sulfates, ions are formed as follows:

$$H_2SO_4 \longrightarrow H^+,\ H^+ + SO_4^{--}$$
$$Na_2SO_4 \longrightarrow Na^+,\ Na^+ + SO_4^{--}$$
$$CuSO_4 \longrightarrow Cu^{++} + SO_4^{--}$$

Likewise, barium chloride solution contains the ions Ba^{++} and Cl^{-}, Cl^{-}. Now when a solution of barium chloride is mixed with any solution containing the SO_4^{--} ion, the two ions Ba^{++} and SO_4^{--} unite to form the insoluble $BaSO_4$, which precipitates; hence the reaction proceeds to completion (paragraph 2, p. 179 of text). *The barium chloride test is therefore really a test for the presence of the* SO_4^{--} ion. Since only sulfuric acid and its salts give this ion, however, it is customary to say that it is a test for sulfuric acid and the sulfates.

EXERCISE 46

THE PREPARATION AND PROPERTIES OF HYDROGEN FLUORIDE

Apparatus. Piece of window glass; small lead dish (laboratory outfit).
Materials. 2 or 3 small pieces of paraffin (size of a pea); 3 g. fluorite (CaF_2); sulfuric acid.

Precaution. *Hydrogen fluoride is very corrosive and must not be inhaled; neither must its solution be brought in contact with the skin.*

1. Place some pieces of paraffin on a glass plate and gently warm the plate over a small flame. When the paraffin is melted, tilt the plate about so as to form a uniform layer of the wax. When the wax is cold, scratch your name *through* the wax with a pin (Fig. 70). Place 3 g. of fluorite in a lead dish and add sufficient sulfuric acid to make a paste of it. Cover the dish tightly with the waxed side of the glass plate and set it in the hood for an hour; then scrape off the paraffin and examine the glass. (*a*) Describe the results. (*b*) Write the equations for all the reactions involved.

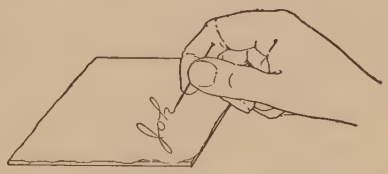

FIG. 70. Preparing a glass plate for etching

EXERCISE 47

THE PREPARATION AND PROPERTIES OF BROMINE

Apparatus. Retort, test tube, and beaker, as shown in Fig. 71; funnel.

Materials. 3 g. sodium bromide or potassium bromide; 4 g. manganese dioxide; 10 cc. sulfuric acid dissolved in 40 cc. water; strips of colored calico; 1 cc. carbon tetrachloride; 1 cc. of chlorine water (R. S.).

Precaution. *The vapor of bromine must not be inhaled.*

1. Preparation of bromine. Put into the retort (*A*, Fig. 71) a mixture of 2 g. of potassium bromide or of sodium bromide and 4 g. of manganese dioxide, and add to this through a funnel the cold dilute solution of sulfuric acid prepared as directed above. Shake the retort so as to mix the contents thoroughly. Note that the test-tube receiver *C* contains sufficient water to allow the end of the retort to dip just below its surface.

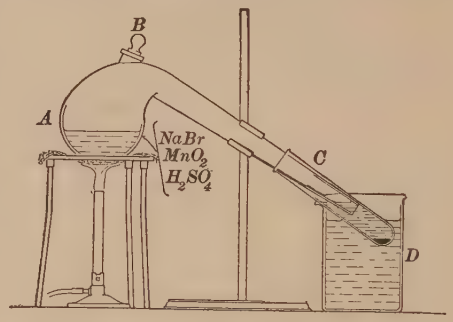

FIG. 71. Preparing bromine by the action of sulfuric acid on a mixture of manganese dioxide and sodium bromide

Now heat the retort gently. The bromine is liberated and distills over. (*a*) Write the equation for the reaction. Continue the heating until all the bromine has distilled over. Remove the stopper from the retort before the heat is withdrawn.

2. Properties of bromine. (*b*) Note the properties of the bromine collected in the bottom of the test tube. (*c*) Has any dissolved in the water (note the color of the water)? (*d*) What property is implied in the name of the element? Add about 0.5 cc. of the aqueous bromine solution to 1 cc. of carbon

tetrachloride. Note that the bromine solution does not mix with the carbon tetrachloride but floats on top of it. Now place your thumb over the mouth of the tube and shake the mixture vigorously for a few seconds; then set the tube aside until the two liquids separate. The bromine, being more soluble in carbon tetrachloride than in water, is taken up by the former liquid. (*e*) What is the color of the carbon tetrachloride solution?

Test the bleaching property of bromine by immersing strips of colored cloth in the aqueous solution. (*f*) How does it compare with chlorine as a bleaching agent?

3. **Action of chlorine on sodium bromide.** Dissolve a bit of sodium bromide as large as a grain of wheat in 2 cc. of water. (*g*) What is the color of the solution? Now add to the solution 1 cc. of chlorine water. The chlorine displaces bromine from sodium bromide just as zinc displaces copper from a copper salt (see displacement series, Exercise 35). (*h*) Account for the change in the color of the solution.

4. (*i*) Summarize the properties and conduct of bromine.

EXERCISE 48

THE PREPARATION AND PROPERTIES OF IODINE

Apparatus. Beaker (250-cc.) and evaporating-dish as shown in Fig. 72; 2 test tubes; stirring-rod.

Materials. 4 g. sodium iodide (or potassium iodide); 4 g. powdered manganese dioxide; sulfuric acid; 2 cc. alcohol; 2 cc. carbon tetrachloride; 10 cc. chlorine water (R.S.); 10 cc. starch solution (R.S.).

1. **Preparation of iodine.** Intimately mix on paper 3 g. of powdered sodium iodide and 2 g. of manganese dioxide. Place the mixture in the bottom of the beaker *A* (Fig. 72). Nearly fill the evaporating-dish *B* with *cold* water. Now remove the evaporating-dish, add 5 cc. of sulfuric acid to the mixture in the beaker, and stir it through the mixture; then at once place the evaporating-dish back on the beaker as shown in the figure. Apply a very gentle heat to the mixture in the beaker, moving

the burner about so as to heat the mixture uniformly. Note the colored vapor formed, which slowly condenses on the sides of the beaker and the bottom of the evaporating-dish in the form of grayish-black crystals. Finally withdraw the flame and let the beaker stand until the vapor is nearly all condensed. (*a*) What is the substance? (*b*) Account for its formation. (*c*) What property does the name of the substance suggest?

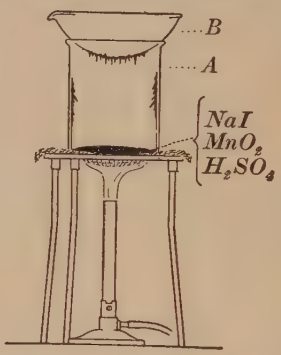

Fig. 72. Preparing iodine by the action of sulfuric acid on a mixture of sodium iodide and sulfuric acid

2. Properties of iodine. By means of a glass rod transfer a little of the iodine collected on the sides of the beaker in 1 to a test tube half filled with water, and shake it for one or two minutes. (*d*) Is the iodine soluble in water (note the color of the water) to any extent? (Set the tube and contents aside for use in 4, below.) Repeat, using alcohol in place of water. (*e*) Compare the solubility of the iodine in water and in alcohol. (*f*) What is the solution in alcohol called? Dissolve a small bit of the iodine in 2 cc. of carbon tetrachloride. (*g*) Note the color of the solution.

3. Action of chlorine on sodium iodide. Dissolve a bit of sodium iodide as large as a grain of wheat in 2 cc. of water in a test tube. (*h*) What is the color of the solution? Now add to the solution 1 cc. of chlorine water. (*i*) Account for the change in the color of the solution.

4. Action of iodine on starch. Place two test tubes in your test-tube rack and label them No. 1 and No. 2. Half fill each with starch solution. To No. 1 add a few drops of the aqueous solution of iodine reserved in experiment 2, above. To No. 2 add a bit of sodium iodide about as large as a grain of wheat and shake the mixture until the iodide dissolves. (*j*) Does free iodine change the color of starch solution? (*k*) Does

iodine in combination change the color of starch solution? Now add 1 or 2 cc. of chlorine water to tube No. 2 and shake the mixture. (*l*) Note and account for the results. (If the color is very deep, pour out a portion of the liquid and replace it with water.)

5. (*m*) Summarize the properties and conduct of iodine.

EXERCISE 49

THE COMPOUNDS OF THE HALOGENS (CHLORINE, BROMINE, AND IODINE) WITH HYDROGEN

Apparatus. 3 test tubes.

Materials. A solution of sulfuric acid prepared by adding (*care*) 10 cc. of the concentrated acid to 2 cc. of water; 1 g. each of sodium chloride, sodium bromide, and sodium iodide; 3 strips of blue litmus paper.

1. Place three test tubes in your rack and label them No. 1, No. 2, No. 3, respectively. In No. 1 put about 1 g. of sodium chloride, in No. 2 a like weight of sodium bromide, and in No. 3 the same weight of sodium iodide. Now add to No 1, drop by drop, 2 cc. of the sulfuric acid, shaking the tube after each addition and waiting for the reaction to subside. Note that a colorless gas is evolved which escapes from the tube and attracts moisture from the air, forming a mist (*this is more noticeable if you blow gently across the mouth of the tube*). Moisten a strip of blue litmus paper with water and hold it at the mouth of the tube. (*a*) What is the gas evolved (Exercise 29)? (*b*) Write the equation for its formation.

2. In a similar way add 2 cc. of the sulfuric acid to tube No. 2, as directed above. Note the colorless gas evolved (blow gently across the mouth of the tube). (*c*) Test it with blue litmus paper as before. (*d*) What is the gas? (*e*) Write the equation for its formation. (*f*) Is a colored substance also formed in the reaction? (*g*) Recalling that hydrogen bromide is unstable, account for the formation of the colored substance.

3. Now add 2 cc. of the sulfuric acid to tube No. 3. (*h*) Is a colorless gas formed as before (blow across the mouth of the tube)? (*i*) Test with blue litmus paper as directed above. (*j*) What is the colored product formed (hydrogen iodide is even less stable than hydrogen bromide)? (*k*) Do you recognize (odor) any gas evolved in addition to hydrogen iodide? (*l*) Account for the formation of this gas, recalling that hydrogen iodide is unstable and that nascent hydrogen is a strong reducing agent.

4. (*m*) Why is the litmus paper used in the above experiments moistened before using? (*n*) Contrast in properties hydrogen chloride, hydrogen bromide, and hydrogen iodide. (*o*) Give the names of the aqueous solution of each.

EXERCISE 50

THE SALTS OF THE BINARY ACIDS OF THE HALOGENS: CHLORIDES, BROMIDES, AND IODIDES

Apparatus. 3 test tubes; mortar and pestle.

Materials. 1 g. each of sodium chloride, sodium bromide, sodium iodide; 1 cc. silver nitrate solution (R.S.); nitric acid; 10 cc. chlorine water (R.S.); 5 cc. carbon tetrachloride; 2 g. manganese dioxide; sulfuric acid.

1. Test for chlorides, bromides, and iodides. Place three test tubes in your rack and label them No. 1, No. 2, and No. 3, respectively. Put into No. 1 a bit of sodium chloride as large as a grain of wheat, in No. 2 put a like amount of sodium bromide, and in No. 3 the same amount of sodium iodide. Now add to each tube about 2 cc. of water and shake the mixtures until the solids are dissolved; then add to each tube 2 or 3 drops of silver nitrate solution. (*a*) Describe the results. (*b*) Write the reaction in each case. Add to each tube 4 or 5 drops of nitric acid and shake the mixture. (*c*) Does the precipitate dissolve? The reaction with silver nitrate serves to detect the presence of chlorides, bromides, and iodides.

2. How to distinguish between chlorides, bromides, and iodides. Arrange three test tubes as directed in experiment 1, above. Thoroughly mix in a mortar a bit of sodium chloride as large as a pea with an equal amount of manganese dioxide and transfer the mixture to tube No. 1. Put in tube No. 2 a similar mixture of sodium bromide and manganese dioxide, and in tube No. 3 a like mixture of sodium iodide and manganese dioxide. Set the tubes in the hood and add to each 4 or 5 drops of sulfuric acid. If the reaction is not marked, heat the tubes gently. (*d*) How can you distinguish between chlorides, bromides, and iodides?

Again arrange three clean test tubes in your rack. Prepare in No. 1 a solution of sodium chloride as directed in experiment 1, above; in a similar way prepare in tubes No. 2 and No. 3 a solution of sodium bromide and of sodium iodide respectively. Now add to each solution about 1 cc. of carbon tetrachloride (note that the tetrachloride does not mix with the water but sinks to the bottom of the tube). Shake the mixtures and then set them in the rack for a few seconds until the liquids separate. (*e*) Note the color of the carbon tetrachloride. Now add to each tube about 2 cc. of chlorine water and again shake the contents of the tubes vigorously; then set the tubes aside until the liquids again separate. (*f*) Now note the color of the carbon tetrachloride. (*g*) What is the function of the chlorine water? (*h*) Describe in detail the method for distinguishing between chlorides, bromides, and iodides by using chlorine water and carbon tetrachloride. Note that the silver nitrate test will enable you to tell whether or not a given compound belongs to the group composed of chlorides, bromides, and iodides. Either the test with manganese dioxide and sulfuric acid or with chlorine water and carbon tetrachloride will enable you to distinguish the three members of the group from each other.

3. Obtain from your instructor an unknown sample of a chloride, bromide, or iodide and determine its identity.

EXERCISE 51

THE PROPERTIES OF GASOLINE AND KEROSENE; ACETYLENE

Apparatus. 2 test tubes; 250-cc. wide-mouthed bottle; medicine dropper or pipette.

Materials. Filter paper; 1 cc. cottonseed oil; wooden splint; 3 cc. gasoline; 3 cc. kerosene; a bit of calcium carbide (CaC_2) the size of a bean.

1. Properties of gasoline and kerosene. (*Gasoline is easily ignited. The student must keep it away from all flames.*) Draw two circles (as large as a silver dollar) on a filter paper. Mark one circle *G* and the other *K*. Within the circle marked *G* place 1 drop of gasoline (use a medicine dropper) and within the circle marked *K* place 1 drop of kerosene. Suspend the paper so that both sides of the paper are exposed to the air, and note from time to time. (*a*) Which product is the more volatile?

Pour about 2 cc. of gasoline into a test tube and add 2 drops of some fat or oil such as cottonseed oil. (*b*) Does the oil dissolve in the gasoline? (*c*) What use does this suggest for gasoline?

Pour about 1 cc. of gasoline into a test tube and add a like volume of water. (*d*) Do the two liquids mix? (*e*) Which is the heavier? (*f*) Repeat, using kerosene. (*g*) Can burning gasoline be extinguished by water?

Gently warm a 250-cc. wide-mouthed bottle by rotating it *above* a flame. Add to the warm bottle 2 or 3 drops of gasoline, place the palm of your hand tightly over the mouth of the bottle, and shake the bottle vigorously. Remove your hand and, standing at arm's length, bring a lighted splint to the mouth of the bottle. (*h*) Record the results. (*i*) What application of gasoline is based on the property noted in this experiment?

2. **Acetylene.** Nearly fill a test tube with water and drop into it a small piece of calcium carbide. (*j*) What is the gas evolved? (*k*) Write the equation for its preparation. (*l*) Ascertain by holding a lighted splint at the mouth of the tube whether the gas is inflammable.

EXERCISE 52

SOME DERIVATIVES OF METHANE: CHLOROFORM, CARBON TETRACHLORIDE, IODOFORM

Apparatus. Test tubes; beaker (100-cc.); glass rod.

Materials. 2 cc. chloroform; 1 cc. cottonseed oil; 3 cc. carbon tetrachloride; splint; 2 cc. alcohol; 5 cc. iodine solution (R. S.); sodium hydroxide solution.

1. **Chloroform.** (*a*) Note its odor. Add about 1 cc. of chloroform to an equal volume of water. (*b*) Does it mix with water? (*c*) Is it heavier or lighter than water? Add 2 drops of an oil (such as cottonseed oil) to 2 cc. of chloroform. (*d*) Is chloroform a good solvent for oils? (*e*) Determine whether chloroform will burn by dipping the end of a glass rod into the liquid and then bringing it in contact with the edge of a flame for an instant.

2. **Carbon tetrachloride.** (*f*) Repeat all the experiments listed under chloroform, substituting carbon tetrachloride for the chloroform. (*g*) Pour about 2 cc. of carbon tetrachloride into a small beaker (100 cc.) and heat it until the vapor of the liquid fills the beaker; then thrust a lighted splint into the vapor. Mention two important commercial uses of carbon tetrachloride illustrated in the above experiments. (*h*) What is the great advantage of carbon tetrachloride over gasoline as an agent in cleaning clothes?

3. **Iodoform.** Pour about 2 cc. of alcohol into a small beaker and add 5 cc. of iodine solution. Now add to the mixture a solution of sodium hydroxide, drop by drop, stirring the liquid after each addition. Continue until the color of the iodine is

completely discharged. Note the formation of a yellowish solid (iodoform). (*i*) Note its odor. (*j*) What is the chief use of the compound?

EXERCISE 53

A STUDY OF FLAMES

Apparatus. Wire gauze; porcelain dish; blowpipe.

Materials. Charcoal (size of a bean); 5 cc. limewater (R. S.); candle; wooden splint; piece of charcoal 8 × 3 × 2 cm.; bit of lead as large as a pea; 2 g. lead oxide.

1. Parts of a flame. (*a*) Note and account for the difference between the combustion of a wooden splint and that of a piece of charcoal. (*b*) What are the conditions necessary for the production of a flame? (*c*) Light a candle and place it so that the flame is against a black background and is not disturbed by air drafts; then note the different cones in the flame. Test the relative temperatures of different parts of the flame by means of narrow strips of splints. (*d*) Draw a diagram in your notebook showing the different parts of the flame and noting the color of each of the cones. Extinguish the candle flame and hold a lighted splint 2 or 3 cm. from the wick in the little column of smoke rising from the wick. (*e*) Note and account for the result.

FIG. 73. Using a blowpipe to form reducing and oxidizing flames

2. Products of combustion. (*f*) What two elements constitute the main percentage of ordinary fuels? (*g*) What products form when these elements burn in air or oxygen? (*h*) Devise and describe simple experiments to show the presence of these products in the gases evolved by the burning candle.

3. Kindling temperature. (*i*) What is meant by the kindling temperature of gases? When a lamp is first lighted, a film of liquid often spreads over the chimney for an instant. (*j*) Explain. Press a piece of wire gauze halfway down on a Bunsen flame. Notice that the flame does not extend above the gauze. (*k*) Is this due to the absence there of gases that are combustible (test for their presence by means of a lighted splint)?

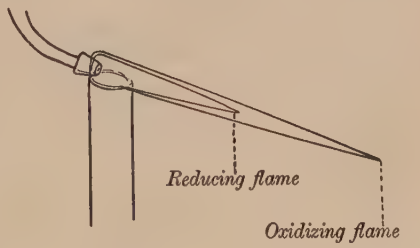

FIG. 74. Showing the location of the reducing and the oxidizing flames

Turn off the gas, then turn it on and ignite it over a piece of wire gauze held horizontally 4 or 5 cm. above the top of the burner. (*l*) Note the results and explain. (*m*) How does the miner's safety lamp prevent explosions?

4. Luminous flames. Hold a porcelain dish in a small luminous Bunsen flame. (*n*) Account for the deposition of carbon. (*o*) Does the nonluminous flame deposit carbon? (*p*) To what is the luminosity of the flame due? Recall the experiment on the Bunsen flame in Exercise 2.

5. Use of the blowpipe. Close the openings on the Bunsen burner so as to secure a luminous flame; then lower the flame until about 3 cm. high. Now place the end of the blowpipe just inside the lower part of the flame and blow gently through the blowpipe (Figs. 73, 74). The flame will be bent over, as shown in the figures. Practice this until you can produce a *continuous* and *steady* flame by forcing a portion of the exhaled

air through the blowpipe by constant tension of the muscles of the cheek. It will be noted that the flame so produced consists of two cones, one inside the other. The inner cone being rich in carbon monoxide is a reducing flame; the outer cone being very hot and in contact with oxygen of the air is an oxidizing flame.

Make a shallow cavity near the end of a piece of charcoal and place in this a bit of lead. Now heat this in the oxidizing flame as shown in Fig. 73. The lead is changed to lead oxide (PbO), a yellow solid which collects about the cavity on the surface of the charcoal.

Clean the charcoal and place in the cavity a paste of lead oxide (PbO) made by adding a little water to 2 g. of the oxide and mixing. Now heat this in the reducing flame (hold the charcoal so that the oxide is in the tip of the inner cone). The oxide will be gradually reduced to a bead of metallic lead.

NOTE. Save the charcoal for future experiments, but before returning it to your desk, pour some water into the cavity to extinguish any fire which may be smoldering in the interior of the charcoal; otherwise it is a source of danger from fire. Indeed, to avoid all danger from fire, it would be better if the charcoal used by the students were kept in some noncombustible vessel such as a crock or pneumatic trough.

EXERCISE 54

SOME COAL-TAR COMPOUNDS (OPTIONAL)

Apparatus. Test tubes; 250-cc. flask, fitted with stopper and glass tube as shown in Fig. 75; glass rod; 250-cc. beaker.

Materials. 10 cc. benzene; 2 drops of cottonseed oil; nitric acid; sulfuric acid; 1 or 2 cc. aniline; 5 g. bleaching powder.

NOTE. The student should first start experiment 2; while the liquids are being heated, experiment 1 can be performed.

Caution. *Benzene is very inflammable and must be kept away from all flames.*

1. Benzene (C_6H_6). (*a*) What is the distinction in composition between benzene and benzine? (*b*) Note the odor

of benzene. (*c*) Add about 1 cc. of benzene to an equal volume of water. (*d*) Does it mix with water? (*e*) Is it heavier or lighter than water? Add 2 drops of cottonseed oil to 2 cc. of benzene. (*f*) Is it a good solvent for oils? (*g*) Dip the end of a glass rod into the liquid and touch it to the edge of a flame for an instant. (*h*) Is benzene inflammable?

2. **Nitrobenzene** ($C_6H_5NO_2$). Pour 8 cc. of nitric acid into a 250-cc. flask and add (*care*), drop by drop, with constant motion of the flask so as to mix the liquids, 10 cc. of sulfuric acid, cooling the mixture all the time by allowing cold water to run on the outside of the flask. When the mixture is cold add, drop by drop, 6 cc. of benzene, mixing the liquids thoroughly after each addition of benzene. Now connect the flask with a cork and glass tube (Fig. 75) and heat the mixture *very gently* (not to boiling point). The tube acts as an air condenser to prevent the escape of the benzene. During the heating mix the liquids in the flask every few minutes by gently rotating the flask. Continue the heating for about half an hour; then disconnect the flask, cool it, and pour the contents into a 250-cc. beaker filled with cold water. Stir the mixture with a glass rod, and after the oil (nitrobenzene) settles to the bottom, decant the clear liquid, leaving the oil in the beaker. Fill the beaker again with water, stir, and decant as before. In this way the acids are washed away, leaving the nitrobenzene pure.

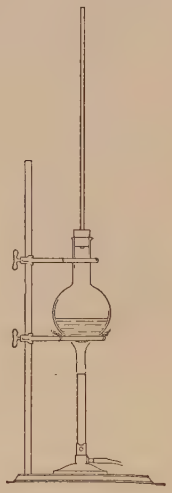

FIG. 75. Preparing nitrobenzene by heating benzene and a mixture of nitric and sulfuric acids

The nitric acid reacts with the benzene to form nitrobenzene and water as shown in the following equation:

$$C_6H_6 + HNO_3 \longrightarrow C_6H_5NO_2 + H_2O$$

The sulfuric acid assists by absorbing the water formed in the reaction. (*i*) Record the color of the nitrobenzene; also its odor (poisonous if too much is inhaled) and density as compared with water.

The above is the first step in the commercial production of *aniline*, which is the source material for many dyes. The nitrobenzene is converted into aniline by reducing it:

$$C_6H_5NO_2 + 3H_2 \longrightarrow C_6H_5NH_2 \text{ (aniline)} + 2H_2O$$

3. Aniline ($C_6H_5NH_2$). (*j*) Note the properties (color, odor, solubility in water) of aniline. Add *one drop* of aniline to 100 cc. of water in a flask and shake the mixture. The aniline soon dissolves in the water. Place about 1 g. of bleaching powder in a beaker, add 25 cc. of water, and stir the mixture thoroughly with a glass rod; then filter and add 1 or 2 cc. of the filtrate to the aniline solution and mix thoroughly. If no change in color takes place, add more of the filtrate, 2 or 3 cc. at a time. (*k*) Record your results. This reaction serves as a delicate color test for aniline.

EXERCISE 55

THE SUGARS

Apparatus. 3 test tubes; 2 small beakers; stirring-rod.

Materials. 10 cc. each of Fehling's solution A and B (R.S.); 5 cc. commercial glucose or corn sirup; 2 g. sucrose; 1 g. each of sweets such as candy, honey, molasses; 2 or 3 g. sodium carbonate dissolved in as little water as possible; red litmus paper; hydrochloric acid.

1. The action of Fehling's solution on dextrose (grape sugar). The most common test for dextrose is the reaction with Fehling's solution. The latter is prepared by mixing equal volumes of solutions A and B, prepared as directed in the Appendix.

Pour into a test tube about 1 cc. each of solutions A and B. When thoroughly mixed, the resulting solution should be deep blue, but perfectly clear. Heat the blue solution nearly to

boiling, add 1 or 2 drops of commercial glucose (corn sirup, procured at any grocery, will do as well), and continue the heating for a few moments. The copper sulfate in the solution is reduced to cuprous oxide by the dextrose, and this separates in the form of a red or yellow solid. Levulose will act in the same way. Dissolve samples of candy, honey, and molasses in a little water and test for the presence of dextrose and levulose in these sweets. (*a*) Note the results obtained.

2. **The action of Fehling's solution on sucrose.** Try the action of Fehling's solution on sucrose as in 1. (*b*) Note the results obtained. Now dissolve about 0.5 g. of the sugar in 5 cc. of water in a test tube and add 3 or 4 drops of concentrated hydrochloric acid. Place the tube in a beaker of boiling water and leave it there for about five minutes; then pour the solution into a small beaker, cool the solution, and neutralize the acid present by adding a concentrated solution of sodium carbonate until the resulting mixture is just alkaline (test with red litmus paper). Now test this with Fehling's solution as in 1. (*c*) Account for the result.

EXERCISE 56

THE COMPOSITION OF MILK

Apparatus. Evaporating-dish; stirring-rods.

Materials. 150-cc. milk; acetic acid (R.S.); Fehling's solution (R.S.); one half junket tablet (these tablets may ordinarily be obtained from any druggist; they may always be had at very little cost by addressing Chr. Hansen's Laboratory, Little Falls, New York).

1. **Determination of the solids and the water in milk.** Place a short stirring-rod in a small evaporating-dish and accurately weigh the two together, recording the weights in the table below. Introduce about 20 cc. of milk and again weigh. Now evaporate the milk to dryness (Fig. 58), occasionally stirring it with the rod to break the scum (protein matter), which would otherwise retard evaporation. While the evaporation is taking place proceed with 2.

When the residue is perfectly dry, carefully wipe the outside of the dish, cool, reweigh, and record the weight. From your results calculate the percentages of solids and water in the milk. Pure milk contains not less than 12 per cent of solid matter. Consult your state laws as to the minimum amount of solid matter allowed in milk sold in the state.

Weight of evaporating-dish and stirring-rod . . .	g.
Weight of evaporating-dish, stirring-rod, and milk .	g.
Weight of milk taken (calculate)	g.
Weight of dish, rod, and residue left after evaporation	g.
Weight of solids in milk taken (calculate)	g.
Percentage of solids in milk taken (calculate) . . .	%
Percentage of water in milk taken (100% — % of solids)	%

2. The separation of the protein (casein) in milk by rennin. Dissolve about one half of a so-called "junket" tablet in a little cold water and add the solution to about 100 cc. of sweet milk heated until just lukewarm. Stir the solution, then cool it and set it aside until the rennin contained in the tablet causes the separation of the curd (casein). Break up the curd by stirring, and filter it. Save the filtrate for further tests. (*a*) For what is the curd used?

3. The separation of the protein in milk by acid. Add 1 or 2 drops of acetic acid (vinegar will serve) to 5 cc. of milk, mix thoroughly, and set aside. The presence of an acid causes the casein to separate (hence the appearance of sour milk).

4. The separation of the lactose in milk. Evaporate the filtrate reserved in 2, refiltering it if more solid matter separates. (*b*) Taste the residue. (*c*) Test it with Fehling's solution (Exercise 55).

5. (*d*) Enumerate the different constituents you have found in the milk.

EXERCISE 57

THE DETERMINATION OF THE FAT PRESENT IN MILK
(Optional)

Apparatus. A Babcock milk tester (Fig. 76) consisting of a centrifugal machine (*A*), 2 test bottles (*B*), pipette made to deliver 17.6 cc. (*C*), and a small cylinder (*D*) graduated to hold 17.5 cc.; beaker (500-cc.).

Materials. Pint bottle of the milk to be tested; sulfuric acid (concentrated, density 1.82 to 1.83; the ordinary commercial sulfuric acid serves well); 100 cc. boiling water.

Note. The most highly prized content of milk is the fat. This is sometimes in part removed by unscrupulous dealers and sold for cream or used in making butter; or water is added to the milk, which reduces

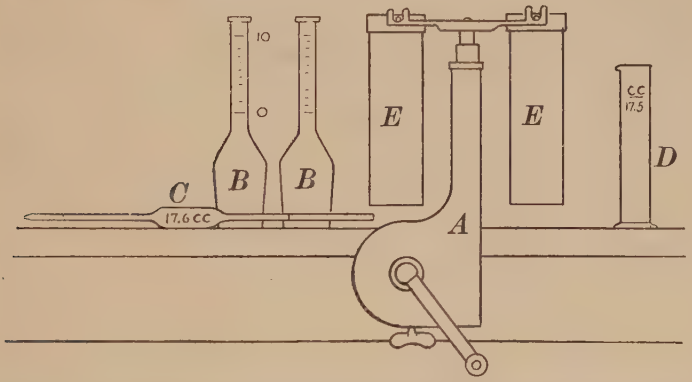

Fig. 76. The apparatus used in determining the percentage of fat in milk by the Babcock method

the percentage of fat. To protect the public, laws have been passed by the different states as well as by the Federal government fixing the minimum percentage of fat in milk sold on the market. Normal milk contains from 3 per cent to 5 per cent of fat, averaging about 4 per cent. The laws of most of the states require not less than 3 per cent. The method ordinarily used for determining the fat in milk was devised by Professor Babcock and is known as the *Babcock test.*

In this test a definite volume of milk is mixed with sulfuric acid in a so-called test bottle, the neck of which is graduated so as to read percentages of fat. The heat evolved in the mixing of the two is sufficient to destroy the casein of the milk but leaves the fat. The bottle is then whirled in a centrifugal machine; in this way the fat, being light, is finally drawn into the neck of the bottle, and the amount present can be read off.

In order to balance the centrifugal machine properly it is necessary to have two test bottles, equally filled, opposite each other. Two different samples can therefore be tested at the same time, or in case only one sample is to be tested, both bottles can be filled with samples of the same milk, and the results compared.

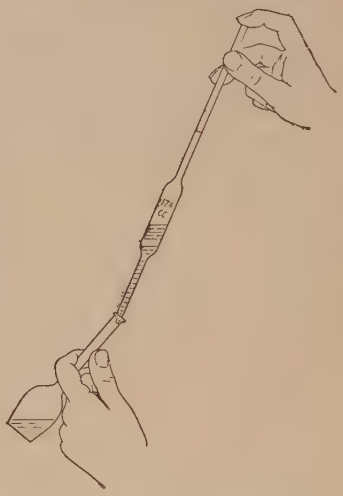

FIG. 77. Showing the proper way of adding milk to the test bottle, determining the fat in milk

1. Introducing the milk and acid into the test bottles. The milk to be tested is first mixed thoroughly by pouring it from the bottle into a beaker, and vice versa, three or four times. By means of the pipette, exactly 17.6 cc. of the milk is then transferred to each of the test bottles, as shown in Fig. 77. Next fill the 17.5 cc. cylinder with the sulfuric acid to the graduation mark; then, holding each of the test bottles in turn at an angle, slowly pour the 17.5 cc. of acid down the wall of the bottle, turning the bottle at the same time so as to wash down the neck of the bottle any milk adhering to the wall. The acid, being heavier than the milk, sinks to the bottom of the bottle.

2. **Mixing the acid and milk.** The milk and acid in each bottle are now carefully mixed by gently rotating the bottle. Considerable heat is developed in the process. Care must be taken not to throw the mixture into the neck of the bottle.

Continue until the acid and milk are thoroughly mixed (the layer of acid at the bottom of the bottle must entirely disappear with the formation of a dark homogeneous liquid).

3. **Separating the fat.** The two bottles are at once placed opposite each other in the containers (*E, E*, Fig. 76) of the centrifugal machine, and the bottles whirled for four or five minutes. The rate of whirling varies with the machine (directions are given with each machine). The machine is then stopped, and hot (nearly boiling) water is added to each bottle until the liquid reaches nearly to the bottom of the narrow neck. The bottles are again whirled for one minute. The machine is again stopped, and hot water is added until the lower part of the fat column reaches about to the 1 or 2 per cent mark on the neck. The bottles are again whirled one or two minutes. The fat should now be clear and should come within the percentage marks on the bottle. By taking the lower and upper readings of the fat column and subtracting, the percentage of the fat in the milk is obtained.

4. (*a*) Give the results of your tests.

EXERCISE 58

A STUDY OF STARCH

Apparatus. Microscope; 200-cc. beaker; stirring-rod; 3 test tubes.

Materials. 1 cc. of iodine solution (R. S.); 0.1 g. flour (largely wheat starch); 10 g. corn starch; piece of bread; hydrochloric acid; 3 g. sodium carbonate dissolved in a little water; red litmus paper; starch solution (R.S.); Fehling's solution (R.S.)

1. **Microscopic appearance.** Examine under the microscope the appearance of starch from different sources (corn, wheat) when magnified. (*a*) Draw diagrams of the starch granules.

2. **Actions of acids on starch.** Try the action of starch solution on Fehling's solution (as in Exercise 55). (*b*) Note your results.

Add 2 cc. of hydrochloric acid to 50 cc. of starch solution in a beaker and boil the contents gently for thirty minutes, allowing the solution to concentrate to about 25 cc. Cool the liquid, neutralize with sodium carbonate, and again test the solution with Fehling's solution. (*c*) Again note the results and explain.

3. Test for starch. Recall the action of iodine on starch (Exercise 48). This constitutes a good test for starch.

Test different foods (such as bread, potatoes, and corn meal) for starch. To do this, boil about 5 g. of the food with 100 cc. of water, stirring the mass thoroughly so as to break it into small pieces; then filter it and cool the filtrate. Now stir the filtrate with a glass rod, the end of which is first dipped into a solution of iodine.

4. The action of heat on starch. Place 3 or 4 g. of starch in a test tube and heat *slightly* for ten or fifteen minutes, stirring it often with a glass rod and regulating the heat so as not to burn the starch (the same results may be obtained by heating a piece of bread in an oven until it is dry and crisp). How does the product differ in taste from the original starch? The heat changes a part of the starch into an isomeric compound known as *dextrin*. (*d*) For what is dextrin used?

EXERCISE 59

TEXTILE FIBERS; PAPER

Apparatus. 4 small beakers or test tubes; stirring-rod; (evaporating-dish; 2 large beakers; microscope; 250-cc. flask fitted with stopper and tube as in Fig. 75).

Materials. 3 strips each of uncolored cotton, pure wool (the kind known as nun's veiling is apt to be pure), silk, and linen cloth (3 cm. × 15 cm.); 50 cc. sodium hydroxide solution; strip of filter paper; hydrochloric acid; sulfuric acid; (3 strips of filter paper 2 cm. × 10 cm.; ammonium hydroxide; sodium hydroxide, desk solution; piece of old white cotton cloth).

1. Effect of heat upon textile fibers. Ignite the end of a strip of cotton cloth in a Bunsen flame; then withdraw it from the

flame. (*a*) Note the odor of the burning cloth. (*b*) Does the cloth when ignited continue to burn?

Repeat, using strips of wool, silk, and linen. (*c*) Can you distinguish in this way between vegetable fibers (cotton, linen) and animal fibers (wool, silk)?

2. How to distinguish between vegetable fibers (cotton, linen) and animal fibers (wool, silk). Place a strip of each kind of cloth in small beakers, cover the cloth with sodium hydroxide solution, and boil the liquid for ten minutes, replacing the water as it evaporates (the cloth must always be completely covered with the liquid); then set the beakers aside until cool. (*d*) Record the results. Cloth sold as woolen often contains considerable cotton. (*e*) How could you detect this adulteration? It will be of interest to test samples of woolen cloth to see if they are pure.

3. How to distinguish between the animal fibers. Immerse strips of silk and wool in concentrated hydrochloric acid and (*f*) note the change after they have stood a few minutes.

4. How to distinguish between the vegetable fibers. Immerse strips of cotton and linen in concentrated sulfuric acid for two minutes. (*g*) Record the results.

5. Miscroscopic appearance of textile fibers. Examine the appearance of each kind of fiber when magnified. (*h*) Compare with the diagrams given in your text.

6. Parchment paper. Pour 20 cc. of sulfuric acid (*care*) slowly and with constant stirring into a beaker containing 10 cc. of water. Pour the solution into an evaporating-dish and allow to cool. Draw strips of filter paper slowly through the acid and then immerse them in a large beaker of water. Finally, wash the strips in a large beaker of water containing 5 or 10 drops of ammonium hydroxide.

(*i*) When the strips are dry compare their properties with those of the untreated paper.

7. Paper pulp. Cut into short, thin strips a piece of old white cotton cloth. Introduce about 2 g. of these strips into a 250-cc. flask and add about 20 cc. of a solution of the sodium hydroxide on your desk. Insert in the mouth of the flask a stopper through which passes a glass

tube (Fig. 75); then heat the flask until the liquid just boils, for two or three hours. When cool decant the liquid and wash the residue several times with water until it is free from sodium hydroxide. This residue is a form of pulp used in making fine writing paper. By beating the pulp until it is very fine, straining it through a fine-wire sieve, and then pressing it with a hot iron, it is possible to obtain a coarse kind of handmade paper.

EXERCISE 60

THE PREPARATION AND PROPERTIES OF COMMMON ALCOHOL

Apparatus. One 2000-cc. flask (or bottle) connected with tube and bottle, as shown in Fig. 78; test tube; stirring-rod; apparatus shown in Fig. 57; evaporating-dish; distilling apparatus (Fig. 41).

Materials. 200 g. molasses or corn sirup; cake of yeast; 50 cc. limewater (R. S.); 25 cc. alcohol; small amounts (size of a pea) of sugar, starch, salt; 1 cc. cottonseed oil; 5 cc. iodine solution (R. S.); sodium hydroxide solution; soda lime (a mixture of sodium hydroxide and calcium hydroxide) sufficient to fill tube *C* (Fig. 78).

1. Preparation of alcohol. (It is suggested that this experiment be performed by the instructor or by students selected by the instructor; after the alcohol is generated, the liquid may be divided among the different members of the class who will then test for the alcohol as directed in 4.)

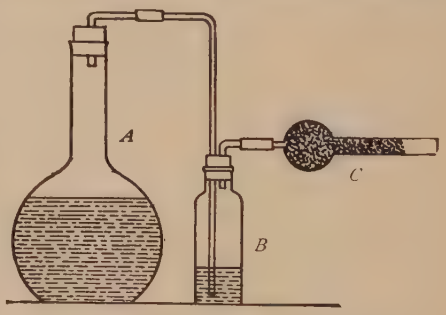

FIG. 78. Preparation of alcohol by the fermentation of molasses

Dissolve about 200 g. of ordinary molasses (or sirup) in 2000 cc. of water in the flask *A* (Fig. 78). Grind a cake of yeast with a little water and add it to the solution in *A*. Connect the flask as shown in the figure (the bottle *B* contains limewater and the tube *C* contains soda

lime). Set the apparatus aside in a warm place (30° is best) for two or three days. Note that a gas is evolved in *A* and bubbles through the limewater in *B*. (*a*) Write the equation for the reaction taking place in *A* and *B*.

2. **Properties of alcohol.** (*b*) Determine the boiling point of alcohol, using the apparatus shown in Fig. 57. (*c*) Pour a few drops of alcohol into an evaporating-dish, ignite, and note the characteristics of the flame.

(*d*) Determine whether alcohol is a good solvent for sugar, salt, starch, and oils, such as cottonseed oil. (*e*) Does alcohol mix with water. (*f*) Contrast the properties of alcohol with those of gasoline.

3. **Test for alcohol.** Pour 2 cc. of alcohol into a test tube and add to this 5 cc. of iodine solution. Now add a solution of sodium hydroxide, *one* drop at a time (mix after the addition of each drop), until the iodine color vanishes; then warm gently and set aside for a few minutes. A yellow precipitate of *iodoform* (Exercise 52), of characteristic odor, forms. (If the amount of alcohol present is small, the iodoform may not separate, but its presence will be revealed by its odor.)

4. When the fermentation of the sugar in *A* (Fig. 78) is complete, divide the liquid so that each student or group of students will have from 150 to 200 cc. Pour the liquid into a flask and distill (Fig. 41) 3 or 4 cc. (*g*) Dip the end of a glass rod in the distillate and touch it to the edge of a flame. Test the remainder of the filtrate for alcohol as in 3 above.

EXERCISE 61

THE COMPOSITION OF FLOUR

Apparatus. Beaker (500-cc.); porcelain crucible; pipestem triangle.
Materials. 50 g. flour; iodine solution (R.S.) for testing for starch; cheesecloth (from 12 to 15 cm. square).

1. Weigh out about 25 g. of flour and mix with just enough water to form a stiff dough, working it in the hands until it

becomes smooth and elastic. Place the dough on a piece of cheesecloth, then fold the cloth about the dough and tie it with a string so as to form a little bag. Nearly fill your largest beaker with water, immerse the bag and contents in the water, and repeatedly squeeze the bag between the fingers. Note that the water becomes cloudy. Retain a portion of this cloudy liquid for tests in 2.

Continue the washing in fresh portions of water until the resulting wash water remains clear. The action may be hastened by working the dough in a small stream of running water. The residue is the nitrogenous constituent of flour and is known as *gluten*. Burn a portion of the gluten and (*a*) note the odor.

2. Test separate portions of the cloudy liquid reserved in 1 for starch (Exercise 58) and for sugar (Exercise 55). (*b*) Record your results.

3. Place about 0.1 g. of flour in a porcelain crucible, heat it with the Bunsen flame (Fig. 30), and gradually increase the heat until only a white residue remains. The residue is mineral matter.

4. (*c*) Enumerate the different constituents you have found in the flour.

EXERCISE 62

THE ACTION OF PRESERVATIVES (Optional)

Apparatus. Two 250-cc. bottles or beakers; 500-cc. beaker; 2 test tubes.

Materials. 300 cc. sweet milk; drop of formalin; 10 cc. hydrochloric acid to which is added *one* drop of a solution of ferric chloride (R. S.); 1 g. sodium benzoate; 100 g. tomato catchup.

1. The action of formaldehyde on milk. Thoroughly clean two small bottles with hot water and half fill each with sweet milk. Add to the milk in one of the bottles *one* drop of formalin and mix thoroughly. Now pour about 3 cc. of milk from each of the two bottles into separate test tubes; add to each an equal volume of the hydrochloric acid to which has been added one

drop of ferric chloride solution (the ordinary commercial hydrochloric acid serves the purpose well, since it usually contains a trace of ferric chloride as an impurity).

Mix the contents of each of the tubes thoroughly and set them in a beaker of boiling water (Fig. 79). Note any change in color. (*a*) How can you detect the presence of formaldehyde in milk? (*b*) What is the distinction between formalin and formaldehyde?

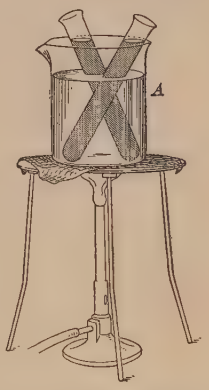

Fig. 79. Heating liquids in test tubes, by placing them in boiling water

Set the two bottles containing the remainder of the milk aside and examine from day to day, noting when the milk in each becomes sour.

2. The action of sodium benzoate on tomato catchup. The student may likewise study the action of sodium benzoate in preventing the fermentation of tomato catchup. Pour the catchup in small beakers, or bottles, adding the benzoate to one of the samples, mixing it thoroughly. Allow the samples to stand exposed to the air and (*c*) note which is the first to sour (odor and taste). In such cases the weight of the preservative added to the catchup should be from 0.1 to 0.2 per cent of the weight of the catchup.

EXERCISE 63

ACETIC ACID: A STUDY OF VINEGAR

Apparatus. Evaporating-dish; beaker (100-cc.); graduated test tube; (burette; graduated pipette; 250-cc. flask).

Materials. 5 cc. acetic acid (R. S.); 30 cc. different kinds of vinegar, such as cider and distilled or white; 1 cc. phenolphthalein solution (R. S.); ordinary desk sodium hydroxide solution; (sodium hydroxide solution of known strength. A so-called *normal* solution serves the purpose well. This contains 40 g. of the hydroxide in 1 l. of solution and may be purchased at any supply house).

1. The properties of acetic acid. Pour about 5 cc. of acetic acid into a small beaker. (*a*) Note its odor. (*b*) Test it with blue litmus paper. (*c*) Add to the acid about 50 cc. of water, mix thoroughly, and taste one drop. Now add to the dilute acid 2 or 3 drops of phenolphthalein solution and then add, drop by drop, the sodium hydroxide solution on your desk, stirring the liquid thoroughly after each addition. Continue until the solution is neutral (see Exercise 31). The reaction is expressed by the following equation:

$$\underset{}{NaOH} + \underset{\text{(acetic acid)}}{H \cdot C_2H_3O_2} \longrightarrow \underset{\text{(sodium acetate)}}{NaC_2H_3O_2} + H_2O$$

The sodium acetate formed is a white solid which may be recovered by evaporating the water. (*d*) To what class of compounds does sodium acetate belong?

2. Determination of the solids in vinegar (quantitative). In this experiment different students should use different kinds of vinegar and compare results. Weigh a small evaporating-dish and record the weight in Table I. Pour into the dish a definite volume (say 25 cc.) of vinegar. Evaporate the vinegar to complete dryness (Fig. 58); then carefully wipe the bottom of the dish with a dry towel and again weigh, recording the weight in the table.

From your results calculate the amount of solid matter in 100 cc. of the vinegar. Pure cider vinegar should contain not less than 1.6 g. of solids in 100 cc. of vinegar (this is the limit fixed by the Federal government as well as by the statutes of many of the states). Distilled vinegar or white vinegar, on the other hand, contains only a trace of solid matter. (*e*) Note the odor and taste of the solid matter obtained from the vinegar. The solids from pure cider vinegar should have an odor and taste suggestive of baked apples. The character of the solids varies according to the source of the vinegar.

TABLE I

Weight of evaporating-dish	g.
Weight of evaporating-dish and residue	g.
Weight of solids (residue) in volume of vinegar taken (calculate)	g.
Weight of solids in 100 cc. of vinegar (calculate) .	g.

3. Determination of the acidity of vinegar. The Federal statutes require that all vinegar sold shall contain not less than 4 g. of acetic acid in 100 cc. of vinegar. The amount of acid present may be determined in the following way:

Prepare or purchase a dilute solution of sodium hydroxide of *known strength*. If a *normal* solution of sodium hydroxide is used, dilute exactly 10 cc. of it to 100 cc. with water and mix thoroughly. 1 cc. of this diluted solution then contains exactly 0.004 g. of sodium hydroxide, which is a convenient strength to use. By means of a graduated pipette (or burette) run into a 250-cc. flask exactly 5 cc. of the vinegar to be tested and dilute with about 50 cc. of water. Add to this 2 drops of phenolphthalein solution. Next fill a burette with the sodium hydroxide solution and run this, drop by drop, into the flask until the solution is neutral, following the directions given in Exercise 32. The liquids must be mixed after each addition of the hydroxide by gently rotating the flask. From the volume of sodium hydroxide solution used, calculate the weight of the hydroxide required to neutralize the acetic acid in the vinegar (see equation in 1, above). From this result calculate the weight of acetic acid present in the 5 cc. of vinegar taken; then from this, the weight of acetic acid in 100 cc. of vinegar. (*f*) Record the results obtained in Table II.

TABLE II

Weight of sodium hydroxide in 1 cc. of the solution used	g.
No. of cc. of hydroxide solution used to neutralize the acid in 5 cc. of vinegar	
Weight of sodium hydroxide required to neutralize the acid in 5 cc. of vinegar (calculate)	g.
Weight of acetic acid in 5 cc. of vinegar (calculate) . .	g.
Weight of acetic acid in 100 cc. of vinegar (calculate)	g.

EXERCISE 64

ESTERS: FATS AND OILS

Apparatus. 250-cc. flask; test tubes; stirring-rod.

Materials. 10 cc. glacial acetic acid (R.S.); sulfuric acid; 10 cc. alcohol; 2 or 3 cc. fats and oils, such as cottonseed oil, olive oil, lard, butter, and cheese; 5 cc. carbon tetrachloride (R.S.); piece of white writing paper.

1. Preparation of a simple ester (ethyl acetate). (*a*) What is an ester? Ethyl acetate is an example of a simple ester. It is a colorless liquid boiling at 78° and has a rather pleasant odor. It is derived from acetic acid ($H \cdot C_2H_3O_2$) by replacing one atom of hydrogen by the univalent radical known as *ethyl* (C_2H_5). This change is easily brought about by heating acetic acid with alcohol, thus:

$$\underset{\text{(acetic acid)}}{H \cdot C_2H_3O_2} + \underset{\text{(alcohol)}}{C_2H_5OH} \longrightarrow \underset{\text{(ethyl acetate)}}{C_2H_5 \cdot C_2H_3O_2} + H_2O$$

Prepare ethyl acetate in the following way: Into a 250-cc. flask pour 10 cc. of glacial acetic acid and add to it (*care*), drop by drop and with constant mixing, 5 cc. of sulfuric acid and then 3 cc. of alcohol. Apply a gentle heat to the mixture. Ethyl acetate is formed and may be recognized by its characteristic fragrant odor. Do not mistake the odor of alcohol for that of ethyl acetate (warm a few drops of alcohol in a test tube and note the difference between its odor and that of the ethyl acetate). The sulfuric acid assists in the reaction because of its affinity for water.

2. Composition and properties of fats and oils. (*b*) To what class of compounds do the fats belong? (*c*) Give the name and valence of the hydrocarbon radical present in fats. (*d*) Give the formulas and names of the three chief constituents of fats and oils. (*e*) What is the main difference in composition between solid fats and oils? (*f*) Test the solubility of typical fats and oils in water, alcohol, and carbon tetrachloride as in Exercise 52. There is no simple accurate test for oils and fats, but the following, based on the fact that fats are nonvolatile, will serve fairly well: In different places on a piece of white writing paper rub one drop of various fats, such as cottonseed oil, cream, melted lard, and butter. (*g*) Note the appearance of the resulting spots when held in front of a light. (*h*) Heat the paper slightly to see if the spots will disappear. (*i*) Test different samples of cheese (warm a small piece and rub it on paper) and milk for fats by this method.

EXERCISE 65

METHODS FOR DISTINGUISHING BETWEEN BUTTER AND OLEOMARGARINE (Optional)

Apparatus. Iron spoon; wooden splint (match stick will do); 100-cc. beaker.

Materials. 10 g. each of fresh butter and oleomargarine (also process butter, if available); 50 cc. sweet milk.

1. Foam test. Over a small flame heat gently in a spoon 2 or 3 g. of the sample (butter or oleomargarine). Under these conditions butter will melt without sputtering and with the formation of much foam on the surface, while oleomargarine will sputter and give but little foam. Process butter, or renovated butter (rancid butter which has been purified by melting the fat, skimming it off the surface, and churning it with milk under certain conditions), acts like oleomargarine. (*a*) Record the results of your tests.

2. Sweet-milk test. Pour 50 cc. of sweet milk into a small beaker and heat nearly to boiling. To the hot milk add 4

or 5 g. of the sample and stir it with a wooden splint until the fat is melted, then place the beaker in ice water and continue the stirring until the fat solidifies. Under these conditions butter will solidify in the form of granules which mix with the milk. Oleomargarine, on the other hand, will collect in a single mass so that it can be removed from the milk in one lump with the wooden stirrer. (*b*) Record your results.

EXERCISE 66

PROTEINS

Apparatus. Evaporating-dish; test tubes; glass rod; 100-cc. beaker.
Materials. Small portion of white of an egg; 25 cc. milk; nitric acid; ammonium hydroxide; 1 g. flour; small strip of woolen cloth.

1. Proteins. Place a small portion of the white of an egg (protein) in a test tube and heat it by dipping the tube into boiling water. (*a*) Record the results.

Pour 20 cc. of milk into an evaporating-dish and heat gently. Note the scum that collects on the surface. Remove the scum with a glass rod, collect other portions of the scum in the same way, and save for further tests.

Transfer a portion of the coagulated white of egg to a small beaker and moisten it with 2 or 3 drops of nitric acid. (*b*) Note any change. (*c*) Wash the egg free from acid with repeated portions of water, then moisten it with ammonium hydroxide and note change in color. These color changes serve as a test for protein matter.

In the same way test for the presence of protein in the scum which separated when the milk was heated. (*d*) Test other substances for protein, such as flour, woolen cloth, a clipping of a finger nail, recording the results. Nitric acid stains the skin yellow; (*e*) suggest an explanation for the change in color.

Burn a small bit of different kinds of protein, such as egg, hair, the scum of milk (note the odor). (*f*) Record your results.

EXERCISE 67

PHOSPHORUS AND ITS COMPOUNDS

Apparatus. 250-cc. wide-mouthed bottle; deflagrating-spoon; glass plate; small beaker; forceps.

Materials. Phosphorus half the size of a pea (to be secured from your instructor when needed); litmus paper (red and blue); 10 cc. ammonium molybdate solution (R. S.); ammonium hydroxide; nitric acid.

1. (*Read carefully the precautions given for handling phosphorus in paragraph 3, Exercise 8.*) Cover the bottom of a wide-mouthed bottle (250-cc.) with water to a depth of about 1 cm. By means of your forceps place the piece of phosphorus on a deflagrating-spoon and ignite it by touching it with a hot wire. Quickly lower the phosphorus into the bottle and cover the mouth of the bottle with a glass plate. When the phosphorus ceases to burn, withdraw the spoon and hold it in the Bunsen flame for a few seconds (to insure that all the phosphorus is burned off the spoon). Cover the bottle with the palm of your hand and shake the water so that the fumes in the bottle are dissolved. (*a*) Test the solution with litmus paper. (*b*) What is present in the water?

2. Pour the solution in the bottle into a small beaker, add 2 or 3 cc. of nitric acid, and boil the solution until about half of it evaporates. The nitric acid oxidizes to phosphoric acid all the phosphorus compounds present. Add a few drops of the solution to 10 cc. of a solution of ammonium molybdate and warm gently. (*c*) Note the result (the compound formed has a very complex composition). Add ammonium hydroxide to the mixture until the liquid is alkaline. (*d*) Note the result. Again acidify the liquid with nitric acid. (*e*) Note the result. *The formation of a yellow precipitate with ammonium molybdate, which precipitate is insoluble in nitric acid and soluble in ammonium hydroxide, serves as a good test for phosphoric acid and its salts.*

EXERCISE 68

SOME COMPOUNDS OF ARSENIC

Apparatus. Hard-glass tube about 12 cm. long and 6 mm. internal diameter, to be made by the student by sealing the end of a piece of hard-glass tubing (*A*, Fig. 80); magnifying-glass; glass rod; iron spoon.

Materials. 0.2 g. arsenious oxide; 2 or 3 g. charcoal in pieces about the size of a grain of wheat.

1. Reduction of arsenious oxide. Into the hard-glass tube introduce an amount of arsenious oxide equal in bulk to about 2 grains of wheat. Cover this to a depth of 2 or 3 cm. with the pieces of charcoal (Fig. 80), which have been previously heated to a high temperature in an iron spoon to expel any volatile matter. See that the inner surface of the tube above the charcoal is perfectly clean.

FIG. 80. Reducing arsenious oxide by hot charcoal

Incline the tube, and heat (use low flame) the *upper* portion of the charcoal to a high temperature (Fig. 80); then, while maintaining the charcoal at this temperature, gradually bring the lower part of the tube also into the flame. The upper part of the tube must not be heated. The arsenious oxide is changed into a vapor, which passes over the hot charcoal and is reduced. The resulting vapor of arsenic condenses on the colder portion of the tube just above the charcoal. (*a*) Note the appearance of the arsenic.

2. Oxidation of arsenic. Cut the tube as near the bottom as possible and remove the charcoal, pushing it out at the bottom of the tube (if necessary) by a glass rod; then, inclining the tube at an angle of about 45°, apply a *very gentle* heat to that portion of it which contains the deposit of arsenic. The air passing through the tube oxidizes the arsenic, forming small white crystals of As_2O_3, which are slowly deposited in the colder portions of the upper part of the tube. (*b*) Examine them with a magnifying-glass. (*c*) Note their form.

EXERCISE 69

A STUDY OF ANTIMONY

Apparatus. Blowpipe; test tube; beaker (200-cc.); stirring-rod; hydrogen sulfide generator (Fig. 68).

Materials. Piece of charcoal (8 cm. × 3 cm. × 2 cm.); 2 pieces of antimony (size of a grain of wheat); hydrochloric acid; nitric acid; strip of zinc half as long as a test tube and narrow enough to go in a test tube; 5 g. ferrous sulfide.

1. Heat a bit of antimony on charcoal in the oxidizing flame (Fig. 73). The product is Sb_2O_3.

2. Into a test tube introduce a bit of antimony no larger than a grain of wheat and add about 3 cc. of hydrochloric acid and then 2 or 3 drops of nitric acid. After the metal has dissolved, pour the solution into 100 cc. of water in a beaker. If a precipitate forms, add hydrochloric acid, drop by drop, with constant stirring, until the precipitate dissolves. Half fill a test tube with the resulting solution and insert a strip of zinc, as in Exercise 35. (*a*) Note the results. Through the remainder of the solution pass a few bubbles of hydrogen sulfide (Exercise 42). The precipitate is Sb_2S_3. (*b*) Note its properties. This is the compound often used in making red rubber.

EXERCISE 70

A STUDY OF BISMUTH

Apparatus. Blowpipe; test tube; beaker (200-cc.); stirring-rod; hydrogen sulfide generator (Fig. 68).

Materials. 2 pieces of bismuth (size of a grain of wheat); piece of charcoal (8 cm. × 3 cm. × 2 cm.); nitric acid; strip of zinc as in Exercise 69; 10 g. ferrous sulfide; hydrochloric acid.

1. Heat a piece of bismuth on charcoal in the oxidizing flame (Fig. 73). Bi_2O_3 is formed and is deposited on the charcoal. (*a*) What is the color of the oxide? (*b*) Write the equation for its formation.

2. Into a test tube introduce a bit of bismuth and add 4 or 5 drops of nitric acid. If the acid does not entirely dissolve the bismuth, add a few drops more. When the bismuth is dissolved, pour the solution into 100 cc. of water in a beaker. If a precipitate forms, add nitric acid, drop by drop, with constant stirring, until the precipitate dissolves. Half fill a test tube with the solution and insert in it a strip of zinc, as in Exercise 69. (*c*) Note the results. Through the remainder of the solution pass a few bubbles of hydrogen sulfide. The precipitate is Bi_2S_3. (*d*) Note its properties.

EXERCISE 71

COMPOUNDS OF SILICON

Apparatus. Evaporating-dish; stirring-rod.

Materials. 2 cc. water glass (solution of Na_2SiO_3); hydrochloric acid.

1. (*a*) Recall the formulas and names of the important acids of silicon.

2. Place 2 cc. of water glass (Na_2SiO_3) in an evaporating-dish, dilute with 10 cc. of water, and add 2 or 3 cc. of hydrochloric acid. Note the gelatinous precipitate of H_2SiO_3. (*b*) Write the equation for its formation. Evaporate the mixture to dryness (Fig. 24) and heat the dish gently with the bare flame. The silicic acid is decomposed into water and silicon dioxide. (*c*) Write the equation for the reaction. When cool, add water, filter, and examine the residue. (*d*) What is it?

3. (*e*) Recall the action of hydrofluoric acid on silica, writing the equation for the reaction.

EXERCISE 72

COMPOUNDS OF BORON

Apparatus. Fine platinum wire (piece 5 cm. long fused in glass tube for handle, as shown in Fig. 81); beaker (100-cc.); stirring-rod; funnel; medicine dropper.

Materials. 8 g. borax; sulfuric acid; 1 or 2 drops of cobalt nitrate solution (R.S.); filter paper.

1. (*a*) Write the names and formulas for three important compounds of boron.

2. Make a little loop on the end of a platinum wire and heat it to redness in a Bunsen flame; then quickly bring the loop in contact with some borax and reheat in the tip of the flame (Fig. 81). The borax adhering to the loop will swell up (owing to the expulsion of the water of hydration) and finally form a clear, glassy bead. (*b*) Note the color imparted to the flame.

Moisten the bead with a drop of sulfuric acid and again touch it to the edge of the flame. (*c*) Note the result. This serves as a simple test for borax.

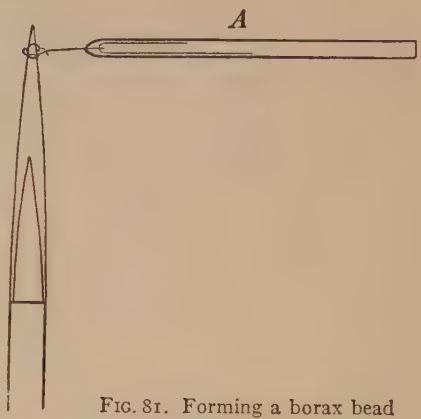

FIG. 81. Forming a borax bead

Moisten the bead with a drop of a solution of cobalt compound and reheat until the bead is transparent when cold. (*d*) Note the color of the bead now (the depth of color depends upon the amount of cobalt present). This property serves as a simple test for cobalt. Some of the other metals likewise impart characteristic colors to the bead, as will be explained later. This method for detecting metals is known as the *borax-bead test.*

3. Dissolve 5 g. of borax in 15 cc. of boiling water. *Carefully* add to the hot solution 2 or 3 cc. of sulfuric acid and stir. The sulfuric acid acts upon the borax as follows:

$$Na_2B_4O_7 + H_2SO_4 + 5H_2O \longrightarrow Na_2SO_4 + 4H_3BO_3$$

The boric acid, H_3BO_3, is insoluble and separates as a precipitate. Filter off the precipitate; then place a bit of it on the loop of the platinum wire used in 2, above, and note the color it imparts to the flame. (*e*) Account for the test for borax given in 2.

EXERCISE 73

COLLOIDS; EMULSIONS

Apparatus. 5 test tubes; filter; 250-cc. flask; funnel.

Materials. Crystal of sodium thiosulfate; hydrochloric acid; filter paper; about 15 g. clay or garden soil; dilute sulfuric acid made by adding 5 drops of sulfuric acid to 5 cc. of water (*care*); solution of aluminium sulfate (R. S.); 2 g. gelatin; white of egg; 5 cc. kerosene; sufficient *thin* shavings of white soap cut from a bar of soap to fill the bottom of a test tube to a depth of about 2 cm.

1. Dissolve a crystal of sodium thiosulfate about the size of a pea in 10 cc. of water and add 1 cc. of dilute hydrochloric acid. Note the separation of finely divided sulfur, which is formed in accordance with the following equation:

$$Na_2S_2O_3 + 2HCl \longrightarrow 2NaCl + H_2O + SO_2 + S$$

(*a*) Is sulfur soluble in water? Filter the mixture. (*b*) Does the sulfur collect on the filter paper? Allow the liquid to stand until the end of the hour. (*c*) Does the sulfur settle? (*d*) In what state is this sulfur?

2. Put a lump of moist clay or garden soil about the size of a small marble in a 250-cc. flask and add about 100 cc. of distilled water. Stopper the flask and shake it vigorously. Allow the coarse material to settle for a few minutes. Pour off equal portions of the muddy water into three tubes marked respectively A, B, and C. Place tube A in the test-tube rack for comparison. To tube B add 1 or 2 cc. of dilute sulfuric acid. Shake it vigorously and set it beside tube A. To tube C add 1 or 2 cc. of a solution of aluminium sulfate and shake the tube vigorously; then set the tube in the rack with the other two. (*e*) Note the rate at which suspended material in the three tubes settles. Any electrolyte will serve in place of the two reagents you have used. (*f*) Why does a river emptying into the ocean fill up at its mouth?

3. Place in a test tube a sufficient amount of small pieces of dry gelatin to make a layer 1 or 2 centimeters in depth. Add about 10 cc. of water and heat the water until a clear liquid is obtained. Now immerse the tube in ice water and note the formation of a gel. Again heat the substance until a clear liquid is obtained, and then repeat the cooling process. (*g*) Does the gel "set" again?

4. Add a small portion of the white of an egg to 10 cc. of cold water and shake the mixture thoroughly. The albumen of the egg is present in the water in the form of a colloid. Divide the mixture into two equal portions. Heat the one portion to boiling. (*h*) Is the colloid precipitated? To the other portion add 2 or 3 drops of nitric acid. (*i*) Note the results. Make suitable test to determine whether the coagulation of the albumen of the egg is a reversible process.

5. Select two test tubes of the same size. Place in the bottom of one to a depth of about 2 cm. some *thin* pieces of white soap cut from a bar of soap; then fill the tube about two thirds full of distilled water. Warm the mixture and shake it gently until the soap dissolves. Now fill the other tube about two thirds full of water. Next pour into each tube about 2 cc. of kerosene. Note that the oil floats on the top of the water in each tube.

Now take a tube in each hand, close the mouth of the tube firmly with your thumb, and shake the tube and contents vigorously for one minute. (*j*) Note the appearance of the liquids in each tube. (*k*) What is such a mixture called? Set the tubes aside five or ten minutes (or until the next laboratory period) and again examine the liquids in the tubes. (*l*) Have the oil and water separated in both tubes? (*m*) What term is applied to a substance, such as soap, that prevents or at least delays the separation of the liquids in an emulsion?

EXERCISE 74

GENERAL METHODS FOR THE PREPARATION OF THE COMPOUNDS OF THE METALS

Apparatus. 6 test tubes.

Materials. 0.1 g. of each of the following salts, dissolved in 5 cc. water: (1) calcium chloride, (2) lead nitrate, (3) barium chloride, (4) ferric chloride, (5) silver nitrate, (6) potassium iodide (solutions of any of these on the Reagent Shelf may be used); sodium carbonate solution (R.S.); hydrochloric acid; sulfuric acid; ammonium hydroxide.

1. By the direct union of the elements. Recall the formation of sulfides of copper and of iron (Exercise 40); of the chloride of antimony (Exercise 28). (*a*) Write the equations for the reactions involved.

2. By dissolving a metal or its hydroxide in appropriate acids. Recall the formation of zinc sulfate (Exercise 9); of sodium chloride (Exercise 31); of copper nitrate (Exercise 37). (*b*) Write the equations for the reactions. (*c*) When a metal or its hydroxide is acted upon by an acid, what becomes of the metal?

3. By acting upon a salt of an acid with an acid having a higher boiling point. Recall the action of sulfuric acid upon sodium nitrate (Exercise 37); of hydrochloric acid upon iron sulfide (Exercise 41); of sulfuric acid upon fluorides (Exercise 46); of sulfuric acid on chlorides (Exercise 29); of hydrochloric acid on carbonates (Exercise 33). (*d*) Write the equations for each reaction and show in what respects they are all similar.

4. By the decomposition of a compound. Recall the action of heat upon potassium chlorate (Exercise 7); upon copper nitrate and lead nitrate (Exercise 38). (*e*) Write the equations for the reactions involved.

5. The following compounds are insoluble (see Appendix, Table of Solubilities of some of the compounds of the metals):

calcium carbonate ($CaCO_3$), lead sulfate ($PbSO_4$), barium carbonate ($BaCO_3$), ferric hydroxide ($Fe(OH)_3$), silver chloride (AgCl), lead iodide (PbI_2). Prepare a small amount of each in a test tube. (*f*) Write the equation for the reaction involved in each case.

EXERCISE 75

THE COMPOUNDS OF SODIUM

Apparatus. Evaporating-dish; beaker; watch glass; stirring-rod; hard-glass test tube; platinum wire; piece of cobalt glass 10 cm. square; magnifying-glass.

Materials. 5 g. sodium carbonate; litmus papers (red and blue); hydrochloric acid; sulfuric acid; 2 or 3 clear crystals of Glauber's salt; 3 g. sodium bicarbonate; 1 cc. limewater (R. S.); 2 g. sodium chloride.

1. Recall experiments with sodium (Exercise 30).

2. Dissolve 5 g. of sodium carbonate in 20 cc. of water. Test the solution with red and with blue litmus paper. (*a*) Account for the results.

Now convert the sodium carbonate present into common salt. (*b*) Describe the method and write the equation for the reaction involved. Treat some of the salt so prepared with sulfuric acid. (*c*) What gas is evolved?

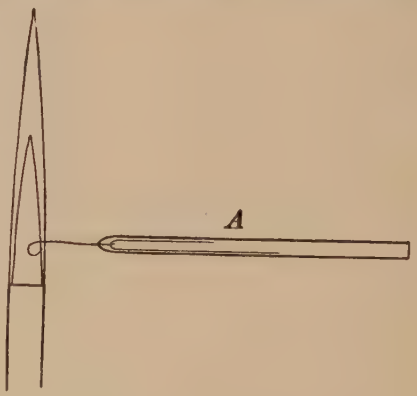

FIG. 82. Method of making a flame test

3. (*d*) Write the complete formula for Glauber's salt. Select some *clear* crystals of the salt, place them on a watch glass, and expose them to the air for one hour or more. (*e*) Note the change. The change is due to the fact that some of the water of hydration escapes. Crystals that behave in this way are said to be *efflorescent*.

4. Fill a hard-glass test tube about one fourth full of sodium bicarbonate and heat gently. Prove that carbon dioxide is evolved (Fig. 54). What liquid condenses in the colder part of the tube? (*f*) Write equations for the reactions by which sodium carbonate is converted into the bicarbonate and vice versa.

5. Bend the end of a platinum wire into the form of a small loop and hold it in the lower part of the outer film of the Bunsen flame (Fig. 82) until it ceases to give any color to the flame; then dip it into a solution of a compound of sodium so that a drop of the solution is suspended in the loop; then hold it in the Bunsen flame as before. (*g*) Note the color. (*h*) Note the appearance of the sodium flame when viewed through a piece of cobalt glass.

6. (*i*) Recall the action of hydrochloric acid on sodium thiosulfate (Exercise 73, 1).

7. Dissolve about 2 g. of sodium chloride in a little water in a beaker and set the uncovered solution aside until the next laboratory period; then (*j*) examine the shape of the crystals with a magnifying-glass. Save the crystals for reference in a future exercise.

EXERCISE 76

THE DETERMINATION OF THE WEIGHT OF COMMON SALT OBTAINED BY ADDING HYDROCHLORIC ACID TO A DEFINITE WEIGHT OF SODIUM BICARBONATE (Optional)

Apparatus. Evaporating-dish and watch-glass cover; medicine dropper; beaker; stirring-rod; balance.

Materials. 1 g. sodium bicarbonate; hydrochloric acid.

1. Carefully weigh to 0.01 g. the evaporating-dish and watch glass, recording the weight in the table below. Transfer to the dish about 1 g. of sodium bicarbonate and reweigh, recording the weights. Pour 4 or 5 cc. of water on the bicarbonate, and place the watch glass on the dish so that only the lip of the dish remains uncovered. Using a medicine dropper or pipette, drop down the lip of the dish 2 or 3 drops of hydrochloric acid.

Wait until the effervescence caused by the escape of the carbon dioxide ceases, then add a few drops more of the acid. Repeat until the addition of the acid no longer causes any effervescence. Now hold the watch glass in the hand just above the dish, and with a little water carefully rinse back into the dish the liquid which has collected on the undersurface of it. Remove the watch glass and slowly evaporate the solution (Fig. 58).

When the solution has evaporated nearly to dryness, cover the dish with the watch glass and heat the dish with the tip of the flame. Continue the heating until there is no more liquid left in the dish or clinging to the undersurface of the glass. Then withdraw the heat and, after the dish is cool, reweigh the dish, watch glass, and residue (sodium chloride).

From your results calculate the amount of salt formed from 1 g. of the bicarbonate, recording the weights. Make the calculations called for in the following table:

Weight of evaporating-dish and watch glass	g.
Weight of evaporating-dish and watch glass plus sodium bicarbonate	g.
Weight of sodium bicarbonate (calculate)	g.
Weight of dish, watch glass, and sodium chloride . .	g.
Weight of sodium chloride (calculate)	g.
Theoretical weight of sodium chloride that can be obtained from the weight of sodium bicarbonate taken	g.

EXERCISE 77

SOME COMPOUNDS OF POTASSIUM

Apparatus. Evaporating-dish; stirring-rod; beaker (250-cc.); cobalt glass; platinum wire; (funnel).

Materials. 10 cc. potassium hydroxide solution (R.S.); hydrochloric acid; 0.5 g. potassium chloride; (17 g. sodium nitrate; 15 g. potassium chloride; filter paper).

1. Pour 10 cc. of a solution of potassium hydroxide into an evaporating-dish, neutralize with hydrochloric acid, and evaporate to dryness. (*a*) What is the product? (*b*) By what other method have you prepared this same compound in a former exercise?

2. Repeat 5, Exercise 75, using a solution of potassium chloride. (*c*) Note the appearance of the flame. (*d*) Also note the appearance of the flame through a piece of cobalt glass. (*e*) What is the difference in the appearance of the sodium flame and of the potassium flame when viewed through the cobalt glass? (*f*) How could you detect both sodium and potassium if they were present in the same solution?

3. Preparation of potassium nitrate. Dissolve 17 g. of sodium nitrate in 15 cc. of boiling water; also 15 g. of potassium chloride in 30 cc. of boiling water. Mix the two solutions in a small beaker and evaporate (stirring the mixture) to about 20 cc.; then quickly filter the hot solution and set the filtrate (Filtrate A) aside until cold.

The reaction between potassium chloride and sodium nitrate is reversible, and the number of grams of each of the four compounds involved which dissolve in 100 g. of water at 15° and 100° is as follows:

	15°	100°		15°	100°
Sodium nitrate . .	84	180	Potassium nitrate .	26	246
Potassium chloride .	33	57	Sodium chloride .	36	40

From a study of these solubilities (*g*) what compound should you expect would separate when the hot solutions of sodium nitrate and potassium chloride are mixed together? (*h*) Examine the crystals on

the filter paper with a magnifying-glass and compare with 7, Exercise 75, to see if your conclusion is correct. (*i*) Taste the crystals.

(*j*) What solid should you expect would separate from Filtrate A when it is cooled? (*k*) Should you expect it to be pure? Filter off the solid and examine it with a magnifying-glass. (*l*) Can you detect crystals of sodium chloride in this solid?

Dissolve the solid in as little *hot* water as possible, cool the solution, and again filter off the solid. Repeat until no crystals of sodium chloride can be detected. Prove the identity of this compound (test for potassium by the flame and for a nitrate by 2, Exercise 38).

EXERCISE 78

THE PROPERTIES OF AMMONIUM COMPOUNDS

Apparatus. Evaporating-dish and watch-glass cover; 5 test tubes.

Materials. Ammonium hydroxide; hydrochloric acid; sodium hydroxide solution; litmus papers (red and blue); (0.2 g. ferrous sulfate solution (R.S.); ammonium carbonate (R.S.); barium chloride (R.S.); calcium chloride (R.S.)).

1. Pour 10 cc. of ammonium hydroxide into an evaporating-dish, neutralize with hydrochloric acid, and evaporate *just* to dryness (Fig. 58). (*a*) What is the residue? (*b*) Note its odor.

Introduce about one half of the residue into a test tube, add a few drops of sodium hydroxide solution, and heat gently. (*c*) Note the odor of the evolved gas and its action on a moist strip of red litmus paper. *All ammonium compounds evolve ammonia when heated with sodium hydroxide. This reaction serves as a good test for all ammonium compounds.*

Cover the evaporating-dish containing the remainder of the residue with a watch glass and heat *gently* with a *small* flame. Note that the solid *sublimes*; that is, passes directly from the solid form into a vapor which condenses (partly) on the cold surface of the watch glass.

2. Add a few drops of ammonium carbonate solution to separate solutions of a compound of barium and of calcium. (*d*) Write the equation for the reaction in each case.

3. Add 2 or 3 drops of ammonium sulfide to 5 cc. of a solution of ferrous sulfate. (*e*) Note the result and write the equation for the reaction that takes place. Repeat, using hydrogen sulfide solution in place of ammonium sulfide. (*f*) Account for the difference in the reaction (consult Table of Solubilities in Appendix).

EXERCISE 79

DETECTION OF COMPOUNDS OF THE ALKALI METALS (Optional)

1. Recall such reactions of sodium, potassium, and ammonium compounds; also of carbonates, sulfates, nitrates, chlorides, bromides, iodides, and phosphates as will serve to identify them, and outline a method of procedure for the identification of them. Then ask the instructor for unknown compounds falling within this list and identify them, recording your results.

EXERCISE 80

THE PREPARATION AND PROPERTIES OF SOAP

Apparatus. Evaporating-dish; large beaker; stirring-rod; funnel; small beaker; 4 test tubes.

Materials. 10 cc. alcohol; 5 g. cottonseed oil (or lard); 1 g. sodium hydroxide dissolved in 2 cc. water; filter paper; hydrochloric acid; magnesium sulfate (R.S.); calcium chloride solution (R.S.); 1 g. sodium chloride dissolved in 5 cc. water.

1. Add 10 cc. of alcohol to 5 g. of cottonseed oil in an evaporating-dish. To the resulting mixture add 1 g. of sodium hydroxide dissolved in 2 cc. of water. Evaporate carefully (use small flame and do not let the tip touch the dish), stirring the mixture constantly until the odor of alcohol can no longer be detected. (*a*) Write the equation for the reaction on the supposition that the oil is composed of olein only. (*b*) What is the process called? (*c*) What remains in the dish?

2. Add 50 cc. of distilled water to the residue in the dish, stir well for a few minutes, and filter, if not clear. Pour a few drops of the filtrate on your hands and rub them (add more distilled

water if necessary) to see if the soap lathers freely. Pour 5 cc. of the filtrate into each of three test tubes. To the first add 2 or 3 drops of hydrochloric acid; to the second add a few drops of a solution of magnesium sulfate. In like manner add a few drops of a solution of calcium chloride to the third. (*d*) Note what takes place in each test tube, writing the equations for each of the reactions. (*e*) Why do waters containing calcium and magnesium compounds (*hard waters*) form a precipitate (curdle) with soap?

3. Add a few drops of sodium chloride solution to 10 cc. of the filtrate obtained in 2. The soap in the filtrate is not in solution but in colloidal suspension. The salt causes the precipitation of the colloid (soap). (*f*) What advantage is taken of this reaction in the manufacture of soap?

4. Recall the effect of soap in the formation of emulsions (Exercise 73). (*g*) What influence has this property upon the cleansing action of soap?

EXERCISE 81

A STUDY OF SOME OF THE COMPOUNDS OF CALCIUM

Apparatus. Blowpipe and charcoal; 2 test tubes; glass tubing for blowing air through solution; evaporating-dish; watch glass; copper penny; piece of window glass; 250-cc. flask; platinum wire for flame test; (spectroscope).

Materials. 3 g. marble; hydrochloric acid; ammonium carbonate (R.S.); drop of cottonseed oil; 10 g. plaster of Paris; 10 g. commercial cyanamide; red litmus paper; 0.1 g. of the chloride or nitrate of each of the metals calcium, strontium, barium (0.1 g. each of sodium chloride and potassium chloride).

1. Place one or two small pieces of marble about as large as a bean on a piece of charcoal and heat them for five minutes in the oxidizing blowpipe flame (Fig. 73). When cool drop the resulting mass into 10 cc. of cold water in a test tube. (*a*) What evidence have you of a chemical reaction taking place? Shake the tube for about one minute; then filter and collect the fil-

trate. Now gently blow exhaled air through the clear filtrate. (*b*) Note all the changes and write the equations for the three reactions that have taken place.

2. Dissolve 1 or 2 g. of marble in hydrochloric acid. (*c*) Write the equation for the reaction. (*d*) What does the effervescence indicate? Evaporate the solution to dryness (use the bare flame and evaporate to complete dryness). (*e*) What is the composition of the residue? Place a small piece of it on a watch glass and expose it to the air for an hour or more (it may be left until the next laboratory period). (*f*) Account for the results.

Dissolve the remainder of the residue in the evaporating-dish in 5 cc. of water (filter if not clear) and add to it a few drops of ammonium carbonate. (*g*) Note the results and write the equation for the reaction involved.

3. Place on a glass plate a penny that has been rubbed with a drop of oil. Pour over the coin a thick paste made by adding a little water to plaster of Paris. Set the glass plate aside until the paste hardens; then remove the coin and (*h*) note the results.

4. Place about one tablespoonful of commercial cyanamide in a 250-cc. flask, add 50 cc. of water, and heat to boiling. Suspend a strip of red litmus paper just inside the flask and note any change in color. (*i*) What do the results indicate? (*j*) Can you detect the odor of ammonia in the escaping steam? Account for its formation. (*k*) What is the original source of the nitrogen present in the ammonia that is evolved (recall the method of making the cyanamide)?

5. Try the flame tests for each of the following metals as in 5, Exercise 75 (use the chloride or nitrate of each metal): calcium, strontium, barium. (*l*) Record the results below.

Metal	Color of Flame
Calcium	
Strontium	
Barium	

6. The use of the spectroscope. The spectroscope is an instrument which has been of great service in many chemical investigations, especially in the discovery of new elements and in detecting the presence of known elements in complex mixtures. The principle involved in its construction is as follows:

When a beam of light passes through a triangular prism of glass, it is bent out of its course and emerges at a decided angle with its original direction, as shown in Fig. 83. Ordinary light is made up of many different wave lengths, and each one is deflected, or *refracted*, to a different degree, so that the various colors of which the light is composed are spread out in a series, the red being the least refracted, the violet the most so. A beam of *white* light gives a continuous series of colors from red through orange, yellow, green, blue, to violet, called a *continuous spectrum*.

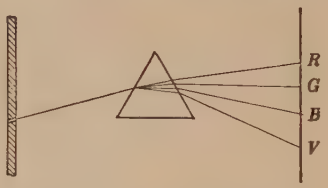

FIG. 83. When passed through a prism a ray of white light is separated into the different colors which together compose the white light

When many of the elements (or their compounds) are volatilized at a high temperature, as in the heat of the Bunsen flame, colored lights are obtained, each metal having its own characteristic color. These lights differ from white light in that they are not made up of so many different wave lengths; hence when passed through the prism the spectrum obtained is not continuous, but merely shows those colors of which the light is composed.

That these colors may be made as distinct and as sharply separated as possible, the light should shine upon the prism through a very narrow slit in a screen, arranged so as to be parallel with the axis of the prism. The colors will then be a series of narrow lines, each an image of the slit, spread out parallel with each other. An instrument, the essential parts of which are a prism, a screen provided with a narrow slit, and lenses for focusing the light upon the slit and for viewing the spectrum, is called a *spectroscope,* or *spectrometer*. Fig. 84 represents a simple form of such a spectroscope, the slit being seen at the end of the tube *B*. The light from the Bunsen burner passes through the slit in the tube *B*, is separated into its different wave lengths by the prism at the center of the spectroscope, and is seen by looking through the tube *A*. The tube *C* contains a scale illuminated by a flame at the end of the tube; this scale makes it possible to fix the position of any lines in the spectrum. Make the following tests with the spectroscope:

First use the luminous Bunsen flame. In this the light is caused by solid particles of carbon, and the spectrum of this will be a continuous band when viewed through *A*. Next examine the spectrum of one or more of the following metals: potassium, sodium, calcium, strontium, barium, using in each case a volatile compound of the metal, preferably the chloride or nitrate. To do this the nonluminous Bunsen flame must be used. First clean the platinum wire until it will impart no color whatever to the flame. Then dip the loop at the end of the wire into a concentrated solution of sodium chloride and bring the loop into the

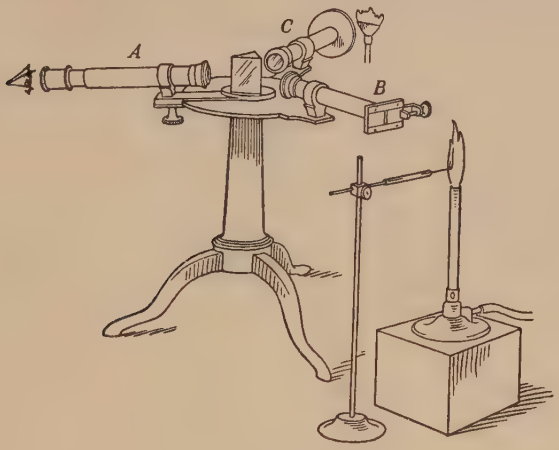

FIG. 84. A spectroscope

lower edge of that part of the Bunsen flame which is next to the tube *B*, as shown in the figure. The spectrum will be found to consist of a single band of yellow light. The slit at the end of the tube *B* should be adjusted by the attached screw until the band is narrow and sharp. In the same way examine the spectrum produced by the other metals mentioned above. It is, of course, essential that the wire be thoroughly cleaned each time; otherwise we will obtain the spectrum of two or more of the metals. It will be found that with the exception of sodium the other metals mentioned will give more than one band of light. (*m*) Diagram the results obtained for each of the metals.

EXERCISE 82

THE PROPERTIES OF BLEACHING POWDER

Apparatus. 2 small beakers; stirring-rod; carbon dioxide generator (Fig. 53); funnel; test tube.

Materials. 25 g. bleaching powder (see that the bleaching powder comes from air-tight cans; otherwise the powder is apt to be worthless); 10 g. marble and hydrochloric acid for generating carbon dioxide; filter paper; narrow strips of different samples of colored calico.

1. The active agent of commercial bleaching powder is the compound $CaOCl_2$. In addition to this compound the commercial powder contains a certain amount of calcium hydroxide and certain other impurities present in the lime from which the bleaching powder was made.

When sulfuric acid is added to bleaching powder both hydrochloric acid and hypochlorous acid (HClO) are formed. These two react as fast as generated, however, to form water and chlorine. The equations for the reactions are as follows:

$$CaOCl_2 + H_2SO_4 \longrightarrow CaSO_4 + HCl + HClO$$
$$HCl + HClO \longrightarrow H_2O + Cl_2$$

A very weak acid like carbonic acid, on the other hand, liberates only hypochlorous acid as follows:

$$2CaOCl_2 + H_2CO_3 \longrightarrow CaCO_3 + CaCl_2 + 2HClO$$

This is the reaction that takes place when bleaching powder is exposed to the air, the carbonic acid being formed from the moisture and carbon dioxide present in the air. Hypochlorous acid is a good bleaching and disinfecting agent. To test its bleaching property, proceed as follows:

Pour 50 cc. of water over 25 g. of bleaching powder in a beaker. Stir the mixture for two or three minutes; then filter. The active agent of the powder, $CaOCl_2$, is now present in the filtrate. Now pass through the filtrate a slow current of carbon

dioxide as long as a precipitate continues to form; then filter. (*a*) What compounds are present in the filtrate? Immerse some strips of colored cloth in the filtrate for a few minutes. (*b*) Record your results.

EXERCISE 83

HARD WATERS AND METHODS FOR SOFTENING THEM

Apparatus. 60-cc. bottle; carbon dioxide generator (Fig. 53); funnel; 2 test tubes; small beaker.

Materials. 30 cc. limewater (R.S.); bit of soap dissolved in water; 10 to 15 g. of marble and hydrochloric acid for generating carbon dioxide; 1 g. magnesium sulfate (R.S.); filter paper; 1 g. sodium carbonate dissolved in as little water as possible; 1 g. powdered calcium sulfate.

1. Bubble carbon dioxide into 25 cc. of limewater diluted with an equal volume of water. Note that a precipitate forms which gradually dissolves as more of the gas is passed through. (*a*) Account for the results, writing the equations for the reactions involved. The resulting liquid is a good sample of a *hard water*.

Add a few drops of a soap solution to 5 cc. of the hard water and (*b*) note the results.

Divide the remainder of the hard water into two parts. Gradually boil the one part. (*c*) Note and account for the result. To the other part add a few drops of clear limewater and mix intimately. (*d*) Again note and account for the result.

2. Shake 1 g. of calcium sulfate with 10 cc. of water in a test tube for two or three minutes; filter, and add to the filtrate 2 or 3 drops of a saturated solution of sodium carbonate. (*e*) Note and account for the results.

All hard waters contain more or less calcium acid carbonate, calcium sulfate, calcium chloride; also the corresponding compounds of magnesium. (The methods used for removing the calcium compounds likewise serve for removing the magnesium compounds.) (*f*) How could such waters be softened on a

large scale? (*g*) Waters softened in this way would contain what compounds in solution?

3. Test some hard waters from wells by the above methods and (*h*) give the results obtained.

EXERCISE 84

TESTING THE ACIDITY OF SOILS

Apparatus. Watch glass; ordinary kitchen knife with thin blade so that it will readily bend; forceps; (250-cc. flask; tripod).

Materials. Cupful of soil to be tested; strips of blue and red litmus papers; (1 g. barium chloride mixed with 0.1 g. zinc sulfide; strip of filter paper (about 6 cm. × 1 cm.) moistened with a solution of lead acetate (R.S.)).

1. Soils under cultivation tend to become acid, and most crops do not thrive in acid soil; hence the custom of adding lime to soils. The following is a simple way of finding out whether or not a given soil is acid:

Collect about one cupful of an average sample of the soil. By means of a clean knife blade mix the soil intimately. Now take one strip each of red and blue litmus paper (use your forceps to avoid any secretion from your fingers from coming in contact with the papers), moisten the papers with distilled water and press them side by side against the concave face of a watch glass. Pour over the strips 2 or 3 tablespoonfuls of the soil; then add distilled water, a little at a time, and gently mix the water through the soil with a clean knife blade (be careful not to displace the red and blue litmus papers) until there is formed a thick paste which will cling to the watch glass when inverted. The glass is now inverted and placed on the desk so that the strips of litmus paper are visible. After fifteen or twenty minutes note if any change in the color of the litmus paper has taken place. (*a*) Record the results of your tests of different soils. (*b*) What precautions must be taken in order to secure reliable results?

2. A method of testing soils largely used at the present time was devised by Professor Truog of the University of Wisconsin and is known as the "Truog test." Proceed as follows:

Collect a sample of soil as in experiment 1 and introduce 12 or 15 g. of the soil into a 250-cc. flask. Next add to this 1 g. of the barium chloride—zinc sulfide mixture (see materials). Pour into the flask 100 cc. of distilled water and shake the flask very gently. Heat the mixture, using a flame that will cause the liquid to boil in from five to seven minutes. Sometimes the liquid froths just when it begins to boil; in case it begins to froth, withdraw the heat for a moment until the frothing subsides. After the liquid in the flask has boiled for just one minute, place across the mouth of the bottle, as shown in Fig. 85, a strip of lead acetate paper and hold it in this position for two minutes, while the boiling continues; then remove the paper and withdraw the heat. If the soil is acid, hydrogen sulfide will be liberated and will act upon the lead acetate, forming black lead sulfide (PbS). The greater the acidity, the greater the amount of lead sulfide formed, and hence the darker the paper. Since the test is very delicate it is essential that care be taken to use distilled water and to see that the materials are pure and that all apparatus used is clean. If there is any doubt as to the purity of the materials, a blank test should be made. To do this proceed just as directed above, except that the sample of soil is omitted.

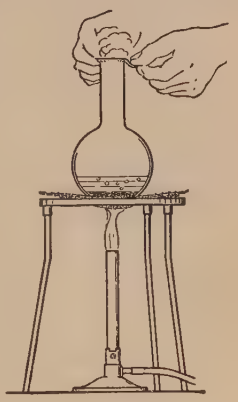

FIG. 85. Testing soils by the Truog method

EXERCISE 85

ACTION OF HARD WATERS ON SOAP (OPTIONAL)

Apparatus. Two 250-cc. bottles; burette or graduated cylinder; 2 test tubes.

Materials. 1 g. soap dissolved in 100 cc. distilled water; samples of hard water (100 cc.); samples of one or more washing-powders.

1. The determination of the amount of soap lost by using hard water for washing. Place two 250-cc. bottles on the desk. Into the first pour 100 cc. of *hard water* (preferably an average

sample of the water used in your town or city) and into the second pour 100 cc. of distilled (or rain) water. Now add to each the soap solution, 1 cc. at a time, and shake the bottle *vigorously* after each addition. Continue adding the soap solution until the lather formed on shaking the mixture persists for five minutes. (*a*) Compare the amounts of the soap solution required in each case to produce a permanent lather. The difference in the amounts represents the soap consumed by using the hard water in place of soft water.

It will be interesting to make a rough approximation of the amount of hard water used yearly in your city for washing, and then to determine approximately the cost of the soap lost in one year owing to the use of hard water.

2. **The analysis of washing-powders.** Devise methods for detecting the presence of the following substances, if present in washing-powders: (*b*) sodium carbonate; (*c*) borax (Exercise 72); (*d*) mineral matter, such as sand. Submit your methods to your instructor for criticism; then test one or more washing-powders for these substances. (*e*) Record your results.

EXERCISE 86

MAGNESIUM AND ITS COMPOUNDS

Apparatus. Porcelain crucible and cover; small beaker; pipestem triangle; evaporating-dish; forceps.

Materials. Strips of magnesium 5 cm. long; blue and red litmus papers; 2 to 3 g. magnesium carbonate; hydrochloric acid.

1. Wind a strip of magnesium wire into a coil and place it in a porcelain crucible. Put the cover on the crucible and apply a gentle heat. By means of your forceps raise the cover slightly from time to time so as to admit air. Continue until the magnesium is entirely burned, leaving a white powder. (*a*) What is the composition of the powder? Add the powder to 25 cc. of water. Stir the mixture and test it with litmus paper. (*b*) Account for the results.

2. Convert 2 or 3 g. of magnesium carbonate into the chloride. (*c*) Describe the process. Evaporate the solution of the chloride to complete dryness in an evaporating-dish, heating the residue gently with the bare flame. When it is cool add a few drops of water, stir, and test with litmus. (*d*) Account for the results. (*e*) Why are waters containing magnesium chloride objectionable for use in steam boilers? (Consult text.)

EXERCISE 87

ZINC AND ITS COMPOUNDS

Apparatus. Blowpipe and charcoal (Fig. 86); 3 test tubes; beaker.
Materials. 0.5 g. zinc; sulfuric acid; sodium hydroxide solution; ammonium sulfide (R.S.); 1 g. zinc sulfate.

1. Place a bit of zinc on charcoal and heat it in the oxidizing flame produced by a blowpipe (Fig. 86). The resulting oxide is deposited as a film on the charcoal. (*a*) Note its color. (*b*) Is its color the same when hot as when cold? (*c*) For what is this oxide used?

Fig. 86. Heating a metal in the oxidizing flame

2. Dissolve 0.5g. of zinc sulfate in 10 cc. of water. Divide this solution into 2 parts and test with the following reagents: (1) sodium hydroxide solution (1 drop, or just sufficient to cause a precipitate); zinc hydroxide precipitates ((*d*) write the equation for its formation), but the precipitate dissolves

again if an excess of sodium hydroxide is added; (2) ammonium sulfide; zinc sulfide precipitates. (*e*) Write the equation for the reaction involved. (*f*) What is the color of zinc sulfide?

3. Perform experiment 1, Exercise 88, so that the exercise may be concluded at your next laboratory period.

EXERCISE 88

RUBBER

Apparatus. 2 beakers (100-cc. and 300-cc.); test tube; stirring-rod; glass plate for covering bottles.

Materials. 1.5 g. natural rubber; 35 cc. carbon tetrachloride; 0.2 g. flowers of sulfur or finely powdered brimstone; acetic acid (R. S.); filter paper.

1. Put about 1 g. of natural rubber in a 100-cc. beaker and pour over it 25 cc. of carbon tetrachloride. Place over the top of the beaker a filter paper, and on this a glass plate so as to prevent the liquid from evaporating. Set the beaker aside until the next laboratory period. Also introduce about 0.1 g. of natural rubber into a test tube and pour over it sufficient carbon tetrachloride to half fill the tube; then cork the tube and set it aside until the next laboratory period (if opportunity is afforded, it is well to stir the mixtures once or twice before the next period). In both cases the rubber forms a colloid gel in the carbon tetrachloride.

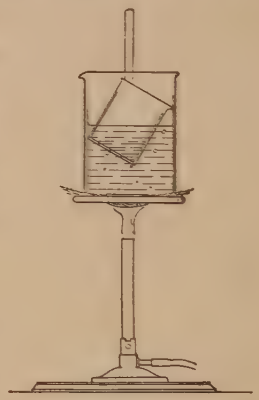

FIG. 87. Evaporating a liquid by placing the beaker containing the liquid inside a larger beaker partially filled with boiling water

2. Stir the mixture in the beaker thoroughly; then add to it about 0.2 g. flowers of sulfur and stir the sulfur until it is uniformly distributed through the mass. Pour into a large

beaker sufficient water so that the smaller beaker containing the rubber will float on the water, as shown in Fig. 87. Heat the water to boiling and maintain the boiling for half an hour or more, until the carbon tetrachloride in the small beaker is all evaporated and the rubber forms a film closely adhering to the sides of the beaker. Then remove the small beaker, and by means of a knife blade peel off the adhering rubber. (*a*) Compare the resulting rubber in properties with the natural rubber.

3. Shake the test tube containing the colloid rubber vigorously so as to form a uniform mass; then add 2 or 3 cc. of acetic acid and again shake the tube. Set the tube aside for a few minutes. (*b*) Does the colloid (rubber) settle? (*c*) What is the action of the acetic acid? (Consult chapter on colloids in text.)

EXERCISE 89

ALUMINUM AND ITS COMPOUNDS

Apparatus. 2 test tubes; blowpipe and charcoal; 2 beakers; funnel.
Materials. 2 g. aluminum turnings; sodium hydroxide solution; wooden splint; hydrochloric acid; ammonium hydroxide; filter paper; aluminum sulfate solution (R. S.); 2 or 3 drops of cobalt nitrate solution (R. S.); 1 g. sodium carbonate in 5 cc. water; (aluminum sulfate and potassium sulfate sufficient to make 20 g. of crystals of potassium alum).

1. (*a*) Note the physical properties of aluminum. Introduce about 0.5 g. of aluminum into a test tube and pour over it 2 cc. of sodium hydroxide solution. Heat the mixture gently. (*b*) What evidence have you that the aluminum is acted upon by the hydroxide? (*c*) Close the mouth of the tube for a few seconds with your thumb; then remove it and *at once* test the gas with a lighted splint. (*d*) What is the gas? (*e*) What would be the effect of washing aluminum cooking vessels with lye or strong soap?

2. Introduce about 1 g. of aluminum into a test tube, add 5 cc. of water and then hydrochloric acid, a drop at a time,

sufficient to dissolve the metal. Filter (if necessary) and dilute the solution to about 50 cc. To this add ammonium hydroxide until the solution reacts alkaline. (*f*) Note the results and write the equation for the reactions involved.

3. Filter off some of the aluminum hydroxide prepared in 2 and heat it on charcoal in the oxidizing flame of the blowpipe (Fig. 86). The aluminum hydroxide is decomposed into aluminum oxide and water. (*g*) Write the equation for the reaction. Now moisten the residue on the charcoal with one or two drops of a solution of cobalt nitrate and again heat strongly in the blowpipe flame. (*h*) Note the change in color. Advantage is sometimes taken of this property in testing for the presence of aluminum.

4. Add a solution of sodium carbonate to any aluminum salt such as aluminum sulfate. Note that a gas is evolved. (*i*) Devise a method for determining whether or not this gas is carbon dioxide and make the test. (*j*) Advantage is taken of this reaction in the preparation of what culinary product?

5. Calculate the weights of aluminum sulfate (the crystals of aluminum sulfate have the formula $Al_2(SO_4)_3 \cdot 18\,H_2O$) and of potassium sulfate required to prepare 20 g. of crystals of potassium alum; then dissolve these amounts of the two compounds separately in as little water as possible, mix the two solutions thoroughly, and set the resulting solution aside for a few days to crystallize. If a string is suspended in the liquid, the crystals will deposit on it. These may then be withdrawn and their properties studied. (*k*) Note the shape of the crystals.

EXERCISE 90

A STUDY OF THE USE OF ALUMINUM SULFATE IN THE PURIFICATION OF WATER

Apparatus. Three 250-cc. wide-mouthed bottles; graduated tube.
Materials. Aluminum sulfate solution (R. S.); limewater (R. S.).

1. Label three 250-cc. wide-mouthed bottles A, B, and C respectively. Nearly fill A and B with muddy water, pouring a like volume of distilled water into C. Add 5 drops of alu-

minum sulfate solution to A and C respectively. Mix the contents of each bottle thoroughly. Now to each of the bottles A and C add 10 cc. of limewater. Set the bottles aside and examine at the beginning of the next laboratory period. (*a*) Record the results and explain the reactions. What is the use of bottle B?

EXERCISE 91

REACTIONS OF BAKING-POWDERS

Apparatus. 250-cc. flask to fit; stirring-rod; mortar and pestle.

Materials. 4 g. sodium bicarbonate; limewater (R. S.); alum; cream of tartar.

1. By referring to the equations for the reactions of different classes of baking-powders as given in your text, calculate and (*a*) record the weight of ammonium alum ($NH_4Al(SO_4)_2 \cdot 12H_2O$) necessary to react with 2 g. of sodium bicarbonate; then grind together these weights of alum and bicarbonate so as to mix them thoroughly. Put the mixture into a 250-cc. flask and cover the mixture with water. Gently rotate the flask so as to form a uniform mixture, at the same time heating the mixture slightly; then set the flask aside for five minutes. Now test for the presence of carbon dioxide in the air in the flask (Fig. 54). (*b*) Note and explain the results.

2. Repeat 1, substituting cream of tartar for the alum. (*c*) Record your results.

3. (*d*) What compounds remain in food as a result of the use of an alum baking-powder; (*e*) of a cream-of-tartar baking-powder?

EXERCISE 92

ANALYSIS OF BAKING-POWDERS (Optional)

Apparatus. 2 small beakers; stirring-rod; funnel; evaporating-dish; 5 test tubes.

Materials. 10 g. each of various kinds of baking-powders; iodine solution (R. S.); filter paper; barium chloride solution (R. S.); hydrochloric acid; sulfuric acid; nitric acid; ammonium molybdate solution (R. S.); sodium hydroxide.

1. Introduce 10 g. of a baking-powder into a beaker and pour over it 50 cc. of water. Stir the mixture thoroughly until no more gas is evolved, then filter it, and test the residue and the filtrate for the various ingredients, as explained below.

2. **Starch.** (*a*) Will any starch present be in the residue or in the filtrate? Make appropriate test. (*b*) Record the test used and the results obtained.

3. **Sulfates.** (*c*) All alums are salts of what acid? (*d*) Make appropriate test for these salts, giving the method and the results.

4. **Tartrates.** Pour 5 cc. of the filtrate into an evaporating-dish, add 5 drops of sulfuric acid, and evaporate to dryness. Finally, heat the dish gently with a bare flame. The presence of a tartrate is indicated by an odor similar to that of burning sugar. (*e*) Record the results of your test of different powders.

5. **Ammonium salts.** Ammonium alum is sometimes used in baking-powders. To detect this, pour 5 cc. of the filtrate into a test tube, add an equal volume of sodium hydroxide solution, and heat gently. If ammonium salts are present, ammonia will be evolved (Exercise 78).

6. **Phosphates.** If calcium or sodium phosphate is present in the baking-powder, the filtrate will contain the acid phosphate of the metal. To detect phosphates, treat 5 cc. of the filtrate with a few drops of nitric acid, heat nearly to boiling, and add a

few drops of the mixture to 5 cc. of ammonium molybdate solution (compare Exercise 67). (*f*) Record the results obtained.

7. Aluminum, calcium, sodium, and potassium. A baking-powder containing sulfates always contains aluminum, while one containing phosphates always contains calcium or sodium; cream-of-tartar baking-powders, on the other hand, always contain potassium. The presence or absence of these metals may be inferred from the tests made for sulfates, phosphates, and cream of tartar.

EXERCISE 93

THE USE OF MORDANTS IN DYEING (Optional)

Apparatus. 200-cc. beaker; stirring-rod; large beaker.

Materials. 2 strips (2 cm. × 6 cm.) of white woolen cloth (nun's veiling serves well); 6 strips of white cotton cloth; 1 g. sodium carbonate in 50 cc. water; 0.5 g. tannic acid in 50 cc. water; 0.2 g. tartar-emetic in 50 cc. water; 0.1 g. of any of the following dyes in 150 cc. water (different students should select different dyes and compare results): fuchsine, methyl violet, gallein, malachite green, Congo red (a solution containing 1 g. sodium carbonate, 5 g. Glauber's salt, and 0.1 g. Congo red in 50 cc. water).

1. Most dyes will dye animal fibers (wool, silk) directly, but will dye vegetable fibers (cotton, linen) fast only when mordants are used.

Place the strips of cotton cloth in a beaker and cover them with the sodium carbonate solution (1 g. in 50 cc. of water) and boil the liquid for five minutes. Remove the strips and thoroughly rinse them with water. This treatment serves to remove all foreign matter from the cloth.

Now *completely immerse* two strips of the cotton cloth in the tannic-acid solution, heat until it is fairly warm to the hand (50° or 60°), and maintain the temperature for ten minutes. Remove the cloth from the solution and squeeze out the liquid, but do not rinse the cloth; then immerse the strips for one minute in the slightly warmed solution of tartar-emetic. The

tannic acid in the cloth reacts with the tartar-emetic, forming a salt (known as antimonyl tannate) which becomes incorporated in the meshes of the fiber and serves as a mordant. Remove the cloth from the solution and rinse it.

Divide the solution of the dye chosen into three equal portions. Heat one portion in a small beaker just to boiling, immerse a strip of *woolen* cloth, and continue the heating for about five minutes. Be sure to keep the cloth entirely immersed in the dye, using the glass stirring-rod for this purpose. Now remove the cloth and rinse it thoroughly. To test whether the cloth is dyed fast, rinse it thoroughly and then wash it in a beaker of water and note whether the water becomes colored. In a similar way dye a strip of *unmordanted* cotton cloth in the second portion of the dye and a strip of *mordanted* cloth in the third portion, and determine in each case whether the cloth is dyed fast. (*a*) Finally, dry the three strips of cloth and insert them in your notebook and record the results of your experiments.

2. Some dyes (known as *substantive dyes*) have the property of dyeing cotton fast without the use of mordants. The ordinary dyes sold by druggists belong to this class. Congo red is a typical dye of this class.

Heat the solution of Congo red, prepared as directed under "Materials," to boiling, immerse in it a strip of *wet* unmordanted cotton, and continue the boiling for five minutes. Remove the cloth, rinse, and test to see whether the cloth is dyed fast. Dry the strip and insert it in your notebook. (The sodium carbonate and the sodium sulfate assist in the process, but do not act as mordants.)

(Save the remaining strips of cotton for use in Exercise 94.)

EXERCISE 94

A STUDY OF LAKES; ALSO THE EFFECT OF USING DIFFERENT MORDANTS WITH THE SAME DYE (Optional)

Apparatus. Two 250-cc. wide-mouthed bottles; evaporating-dish; 2 small beakers; stirring-rods.

Materials. 5 cc. of a 20-per-cent solution of alizarin paste; 0.1 g. of gallein dissolved in 50 cc. water; solutions of aluminum sulfate and ferric sulfate (R. S.); ammonium hydroxide; 3 strips of cotton cloth prepared in Exercise 93.

1. Formation of lakes. Label two wide-mouthed bottles A and B respectively. Shake the alizarin paste so as to form a uniform mixture; then introduce about 10 drops of the paste into each of the bottles. Next add to each of the bottles 2 cc. of ammonium hydroxide and then 200 cc. of water, and mix the contents thoroughly. Now add 10 cc. of aluminum sulfate solution to bottle A and 10 cc. of ferric sulfate to bottle B, intimately mix the solutions, and set the bottles aside; (*a*) note the appearance of the contents at the end of the laboratory period, also at the beginning of the next laboratory period. (*b*) What is the function of each of the materials used?

2. Mordanting strips of cloth with aluminum hydroxide or ferric hydroxide. Pour about 20 cc. of aluminum sulfate solution or ferric sulfate solution into a small beaker and heat to boiling (different students should use different solutions and compare results). Completely immerse in this solution two of the strips of cotton cloth prepared in Exercise 93, and continue the heating for from two to three minutes. Remove the cloth, squeeze it between the fingers to remove the excess of the solution, and immerse it in 20 cc. of water containing from 1 to 2 cc. of ammonium hydroxide. Warm the liquid slightly for two minutes, then remove the cloth and rinse it twice in water. (*c*) What compound is now incorporated in the cloth?

3. Dyeing a strip of mordanted cloth with alizarin. Pour 1 cc. of alizarin paste into a small beaker, add 20 cc. of water, stir, and heat to boiling. Completely immerse one of the strips of mordanted cloth prepared in 2 and continue the heating and stirring for five minutes; then remove the cloth and rinse thoroughly. Dry the dyed strip and insert it in your notebook.

4. Dyeing a strip of mordanted cloth with gallein. Heat the solution of gallein just to boiling; then immerse one of the mordanted strips of cotton in the dye. Keep the strip completely immersed in the dye, and continue the heating for five minutes; then remove the cloth and rinse it. Dry the dyed strip and insert it in your notebook.

EXERCISE 95

THE DETECTION OF DYES IN FOODS (Optional)

Apparatus. Small beaker; stirring-rod.

Materials. Samples of colored pop and orangeade; strips of woolen cloth (nun's veiling gives good results); hydrochloric acid.

1. Select different samples of colored pop. Pour 50 cc. of each into a beaker, add 2 or 3 drops of hydrochloric acid, and heat to boiling; then introduce a strip of woolen cloth and continue the heating for five minutes. Be careful to keep the cloth completely immersed in the liquid. (*a*) Remove the cloth, rinse, and note the color.

2. Samples of colored candies may be tested by first dissolving the candy in water and then testing the solution for dyes by using strips of woolen cloth, as in the above case.

Tomato catchup is sometimes colored, although the practice is forbidden by Federal law. To test a catchup for dyes, heat a portion of the catchup diluted with water and immerse a strip of woolen cloth in the hot mixture for five minutes. Remove the cloth and rinse thoroughly. If artificial dyes are present, the cloth will be deeply colored; otherwise it will have only a slight brownish tinge produced by the natural coloring-matter of the tomato.

EXERCISE 96

CLAY; PORTLAND CEMENT; MORTAR

Apparatus. Evaporating-dish; 2 beakers; a piece of wooden shingle or board about as large as the palm of your hand; ordinary kitchen knife.

Materials. Small samples (half a cupful) of samples of different clays; 30 g. Portland cement; 15 g. sand; 15 g. lime (this should be taken from a lump of lime freshly made); 2 small pasteboard boxes (pill boxes will do).

1. Add water, a little at a time, to different samples of clay and work it intimately through the clay until a thick, plastic mass is obtained. Devise a rough method for testing the relative plasticity of different clays in a general way and apply it to the clays at hand, arranging them in the order of their plasticity. (*a*) Describe the method you used.

2. To 25 or 30 g. of Portland cement add water, a little at a time, and stir it through the mass until a thick paste is obtained. Spread a portion of this on a shingle and set it aside until the next laboratory period. With the remainder fill a small paper pill box. Place the box in a beaker half full of water. Set the beaker and contents aside until the next period. (*b*) Examine the cement placed on the board, also that in the beaker, and note whether the samples have hardened, or "set."

3. Repeat experiment 2, using, in place of the cement, some mortar made as follows: Add to 15 g. of lime 10 cc. of water. The lime will "slake" and form a thin paste with the excess of water. Now work through this, a little at a time, 15 g. of fine sand, adding more water if necessary, to keep the mortar from becoming too thick. Place some of this on a shingle and with the rest fill a paper pill box; then proceed as in experiment 2. (*c*) Describe the results obtained, noting in particular any difference in the properties of Portland cement and mortar.

EXERCISE 97

A STUDY OF IRON AND ITS COMPOUNDS

Apparatus. Forceps; 2 beakers; watch glass; funnel; (flask (250-cc.); 8 test tubes).

Materials. Watch spring from 10 to 15 cm. in length (broken watch springs can be obtained from any jeweler); 5 g. small tacks or fine iron wire; 0.5 g. powdered iron; filter paper; sulfuric acid; (nitric acid; hydrochloric acid; ammonium hydroxide; potassium ferrocyanide (R.S.); potassium sulfocyanate (R.S.)).

1. The tempering of steel. Heat a piece of watch spring (from 10 to 15 cm. in length) to a white heat in a Bunsen flame. Let it cool slowly, and when cold bend it to determine if it is brittle. Again heat to a white heat and *at once* plunge into a beaker of cold water. When cool, bend the piece as before. Reheat the piece, allow it to cool slowly, and again examine it. (*a*) Record the results of the experiments.

2. Preparation of ferrous sulfate (copperas). Place 5 g. of fine iron wire or small tacks in a beaker and pour over it 15 cc. of water. Now add (*care*) 4 cc. of concentrated sulfuric acid and heat very gently (hood) until a vigorous evolution of gas takes place; then cover the beaker with a watch glass and set it aside in the hood until near the end of the laboratory period. Then add 10 cc. of water, and heat slowly until the liquid boils, stirring the mixture constantly. Filter off any undissolved solids, collecting the filtrate in a beaker. Set the uncovered beaker containing the filtrate in your desk until the next laboratory period. (*b*) Record the properties of the crystals. (*c*) Write the equation for their formation from iron.

3. Preparation of ferrous and ferric salts and their reactions. Place about 0.5 g. of iron powder in a small flask and pour over it 5 cc. of water and then 2 cc. of hydrochloric acid. Mix the contents of the flask, heat the flask gently, and set it aside in the hood for about five minutes. The iron dissolves in the hydrochloric acid, forming ferrous

chloride. Now add 50 cc. of water to the flask, mix well, and filter off the undissolved iron. Nearly fill a test tube with the filtrate, add to it 4 or 5 drops of hydrochloric acid and a small piece of iron wire or 1 or 2 tacks, and loosely cork the tube. Mark this "Solution A." It consists of a solution of ferrous chloride. The iron wire and the hydrochloric acid generate a little hydrogen, which prevents the oxidation of the ferrous chloride to ferric chloride by the oxygen of the air. Mark the remainder of the filtrate "Solution B."

To Solution B add 1 cc. of hydrochloric acid and heat it nearly to boiling; then withdraw the flame and add nitric acid, drop by drop, with constant stirring, until the solution, which is at first dark brown in color, becomes light in color (about 2 cc. of nitric acid will be required) The *ferrous chloride* in the solution is changed to *ferric chloride* by the oxygen furnished by the nitric acid, thus:

$$2\,FeCl_2 + 2\,HCl + O \longrightarrow 2\,FeCl_3 + H_2O$$

Now compare the action of the following reagents upon Solutions A and B (add 2 or 3 drops of the reagents to 3 cc. of the solutions in separate test tubes): ammonium hydroxide, potassium ferrocyanide, potassium sulfocyanate (KCNS). Tabulate your results as follows:

	Ferrous Chloride (Solution A)	Ferric Chloride (Solution B)
Ammonium hydroxide		
Potassium ferrocyanide		
Potassium sulfocyanate		

EXERCISE 98

THE REMOVAL OF STAINS

Apparatus. 4 test tubes; 2 beakers; evaporating-dish.

Materials. 0.2 g. tannic acid dissolved in 10 cc. water; ferric sulfate (R. S.); 2 pieces of white cloth 10 cm. square; 1 g. oxalic acid dissolved in 50 cc. water; hydrochloric acid; 2 strips of black cloth 5 cm. square; nitric acid; ammonium hydroxide; a few drops each of cottonseed oil and sirup (molasses); 25 cc. carbon tetrachloride; blotting-paper; strips of cloth stained with coffee and fruit juice; acetic acid (R. S.); 20 g. bleaching-powder; hot water.

1. Stain two strips of white cloth by dipping them into a solution of ferric sulfate until thoroughly saturated and then into

a solution of tannic acid (or they may be stained directly with black ink). Wash one of the strips repeatedly with boiling water. Leave the other exposed to the air until dry; then try the effect of hot water upon the stain. If the stain is not removed, wash with a dilute solution of oxalic acid and finally with hot water. (*a*) Record your results.

2. Place 2 or 3 drops of dilute hydrochloric acid upon a strip of black cloth. (*b*) What color is produced? Wash the spots with 10 cc. of water containing 4 or 5 drops of ammonium hydroxide. (*c*) Do the spots disappear? Repeat, using nitric acid in place of hydrochloric acid. (*d*) Account for the results.

3. Place in separate test tubes 4 or 5 drops of a sirup and a like amount of fat, such as cottonseed oil. Test the solubilities of each in water and in carbon tetrachloride. (*e*) Suggest a method for removing stains made by sirups and one for those made by fats. Stain some strips of cloth with sirup and with an oil and test your methods for removing these stains. (In applying a solvent it is convenient to place the stained portion of the cloth over a piece of blotting-paper. A small bit of sponge or cloth, saturated with the solvent, is then rubbed about the stained portion, gradually nearing the stain itself, which is finally thoroughly rubbed.)

Benzine (or gasoline) may be used in place of the carbon tetrachloride. *If benzine is used, however, it must be remembered that it is very inflammable. Never use it in the vicinity of a flame.*

4. Stain some strips of cloth with coffee and with fruit juices. (*f*) Are the stains removed by washing with boiling water? If the stain cannot be removed in this way, wash the stained portion of the cloth in bleaching-powder to which has been added some water and 3 or 4 drops of acetic acid.

EXERCISE 99

A STUDY OF COPPER AND ITS COMPOUNDS

Apparatus. 3 test tubes; beaker; (porcelain crucible; pipestem triangle; balance).

Materials. Nail; copper sulfate solution (R. S.); sodium hydroxide; ammonium hydroxide; 10 cm. copper wire; hydrochloric acid; (3 g. copper sulfate crystals).

1. (*a*) Recall the action of nitric acid and of sulfuric acid on copper (Exercises 37, 43); also the action of sulfur on copper (Exercise 40). Suspend an iron nail so that half its length is immersed in a solution of copper sulfate for five minutes. (*b*) Account for the result.

2. To 2 cc. of a *cold* solution of copper sulfate in a test tube add one half its volume of sodium hydroxide solution. Copper hydroxide, $Cu(OH)_2$, is precipitated. Now heat to boiling. The hydroxide is decomposed into water and cupric oxide (black). (*c*) Write the equations for the reactions.

3. Add 1 drop of ammonium hydroxide to a dilute solution of copper sulfate; now continue to add the ammonium hydroxide, drop by drop, until the precipitate which is at first formed is dissolved. (*d*) How does the color of this solution compare with that of the original solution? This reaction is characteristic of copper compounds.

4. (*e*) Recall the formation of cuprous oxide (Exercise 55).

5. Moisten the end of a copper wire with hydrochloric acid and hold it in the edge of a Bunsen flame. (*f*) What is the color of the flame? (*g*) What is the function of the acid?

6. The determination of the percentage of water of hydration in copper sulfate (quantitative). Accurately weigh (or counterpoise) a porcelain crucible and cover. Record all weights in the table below. Place 2 or 3 g. of crystals (no larger than a pea) of copper sulfate in the crucible and again accurately weigh. Place the covered crucible on a pipestem triangle and heat it with a gentle flame until the crystals lose

their color. This will require from twenty to thirty minutes. The tip of the flame should not quite touch the crucible. The product is anhydrous copper sulfate. When the crucible is cool, reweigh. From your results calculate the percentage of water of crystallization in the crystals. Fill in the blank spaces in the table below:

Weight of crucible	g.
Weight of crucible plus copper sulfate crystals	g.
Weight of copper sulfate (calculate)	g.
Weight of crucible plus anhydrous copper sulfate	g.
Weight of anhydrous copper sulfate	g.
Weight of water of hydration (calculate)	g.
Percentage of water of hydration in crystals	%
Theoretical percentage of water of hydration	%
Average of results obtained by members of the class	%

EXERCISE 100

A STUDY OF MERCURY AND ITS COMPOUNDS

Apparatus. 100-cc. beaker; 2 test tubes.

Materials. Globule of mercury (size of a grain of wheat); nitric acid; copper penny; 0.5 g. mercuric oxide; 3 cc. solution of mercurous nitrate (R. S.); hydrochloric acid.

1. (*a*) Note the physical properties of mercury. Place a globule of it in a small beaker and add (hood) just enough nitric acid to dissolve it. Dilute the product with 10 cc. of water and place a copper penny in the solution. After a few minutes remove the coin and polish it with a piece of cloth. (*b*) Account for the result (recall the displacement series of the metals).

2. For what purpose have we used mercuric oxide? Place 0.5 g. of it in a test tube and dissolve it in as little nitric acid as possible. (*c*) What compound of mercury is formed? Then add water until the test tube is one fourth full. Into a second test tube pour a similar volume of a solution of mercurous nitrate. Now add 2 or 3 drops of hydrochloric acid to each test tube. (*d*) What conclusions do you draw in reference to the solubility of the two chlorides of mercury?

EXERCISE 101

A STUDY OF SILVER AND ITS COMPOUNDS

Apparatus. 200-cc. beaker; stirring-rod; funnel; blowpipe and piece of charcoal; 3 test tubes.

Materials. Silver dime; nitric acid; hydrochloric acid; ammonium hydroxide; filter paper; hot water; 2 or 3 g. sodium carbonate; 10 cc. silver nitrate solution (R. S.); solutions of potassium bromide and of potassium iodide (R. S.).

1. Place a silver dime in a small beaker and add (hood) sufficient nitric acid to dissolve it. The solution may be hastened by applying a gentle heat. When the solution is complete, dilute the product with about 25 cc. of water. (*a*) Account for the color of the liquid. Now add, drop by drop, with constant stirring, a solution of hydrochloric acid. The precipitate formed settles to the bottom of the beaker when stirred. Continue adding the acid until a drop of the acid no longer causes a precipitate when it comes in contact with the clear liquid in the beaker. Now carefully pour off the clear liquid from the precipitate and add ammonium hydroxide to this liquid until the solution becomes alkaline. (*b*) Account for the change in color (3, Exercise 99). Wash the precipitate remaining in the beaker two or three times by pouring hot water over it and decanting. Finally, remove any remaining water by filtration. Mix the product with an equal bulk of sodium carbonate, transfer to a small cavity in a piece of charcoal, and heat it in the blowpipe flame. The silver salt is gradually reduced to metallic silver, which will fuse into a globule if sufficient heat is applied. (*c*) How does the product differ in composition from that of the original coin?

2. Prepare small amounts of the chloride, the bromide, and the iodide of silver. (*d*) Give the methods you employed. Expose to the sunlight the test tubes containing the precipitates and note any changes. For what are these compounds used?

EXERCISE 102

THE CHEMISTRY OF PHOTOGRAPHY

Apparatus. 4 test tubes; hard-glass test tube; 500-cc. beaker; glass rod; piece of window glass.

Materials. Solution of potassium bromide (R.S.); solution of silver nitrate (R.S.); 0.5 g. silver nitrate; 0.5 g. potassium bromide; 3 g. powdered gelatin; tube of any commercial photographic developer; solution of sodium thiosulfate (R.S.).

1. Preparation of the plate or film. Pour into a test tube 3 cc. of a solution of potassium bromide and add to this 2 cc. of a solution of silver nitrate. (*a*) What is the compound formed? Place the tube and contents in the bright sunlight for two or three minutes. (*b*) Do you observe any change in color?

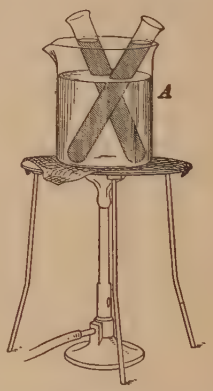

FIG. 88. Heating a liquid in a test tube by immersing the tube in hot water

It is not possible to coat a glass plate or film with the pure silver bromide. The bromide is therefore suspended in a finely divided state in an emulsion of gelatin, and the glass coated with the emulsion. The gelatin keeps the bromide in a finely divided state and also makes it much more sensitive to light. Proceed as follows:

Half fill a 500-cc. beaker with water, heat until the water is uncomfortably hot to your finger, and maintain this temperature with a small flame. Now pour 3 cc. of water into a test tube and add to it 0.5 g. of silver nitrate. Shake the mixture until the nitrate is dissolved; then place the tube in the beaker, as shown in Fig. 88. Next pour 3 cc. of water in your hard-glass test tube (used in preparing oxygen), and place

this in the hot water in the beaker until the water in the tube has about the same temperature as that in the beaker. Next add to the tube 0.5 g. of potassium bromide and then, a little at a time, with constant stirring (use a glass rod), about 2 g. of powdered gelatin. When the gelatin is uniformly distributed through the hot water, add to this, a little at a time, the hot solution of silver nitrate from the tube in the beaker, stirring it thoroughly with a glass rod. The hot emulsion gradually becomes thinner. Now pour this emulsion (keep it away from a bright light) onto a glass plate, tilting the plate until the emulsion forms a uniform coating on the plate. Set this aside for five or ten minutes until the emulsion hardens. In actual practice all these operations must be carried out in a dark room. (*c*) What is the composition of the emulsion? (*d*) Note its color. Now expose the plate to the bright sunlight for a few minutes. (*e*) Note the change in color.

2. The chemistry of the development of image. Pour into a test tube 2 cc. of a silver nitrate solution and add to this 3 cc. of potassium bromide solution. Do not shake the tube, as this will cause the precipitate to form a compact mass. (*f*) Record your results. Expose the tube and contents to the bright light for five or ten seconds; then add to it 5 cc. of a solution of the developer prepared as directed in the instructions accompanying the developer. (*g*) Record and explain the results. Save the tube and contents for experiment 3.

3. The chemistry of "fixing" the negatives. Pour off the liquid from the tube (experiment 2), leaving the solid in the tube, and add to this solid 10 cc. of a solution of sodium thiosulfate (hypo). Shake the mixture vigorously for about one minute. (*h*) Does the precipitate dissolve? Set the tube aside for comparison later.

4. Now repeat experiments 2 and 3 in a dark room (omitting, of course, the exposure to light, as directed in experiment 2). (*i*) Account for the difference in the results obtained when the experiment is performed in the absence of light. (If a dark

room is not available, fairly acceptable results may be obtained by wrapping a piece of black paper around the tube in which the reactions are carried out and performing the experiment in a dark place in the laboratory.)

EXERCISE 103

SOME PROPERTIES OF TIN

Apparatus. 100-cc. beaker; blowpipe and charcoal.

Materials. 2 g. tin; hydrochloric acid; solution of mercuric chloride (R. S.).

1. (*a*) Note the physical properties of tin. (*b*) Heat a bit of it on charcoal (Fig. 86) and note the changes.

2. Dissolve about 1 g. of the metal in hydrochloric acid. (*c*) What is formed? Cool, dilute with a little water, and add 1 to 2 drops of the solution to 3 cc. of mercuric chloride solution. A white precipitate of mercurous chloride forms:

$$SnCl_2 + 2\ HgCl_2 \longrightarrow SnCl_4 + 2\ HgCl$$

Now add a few drops more of the stannous chloride solution and heat the mixture gently. The mercurous chloride is reduced to metallic mercury, which forms a dark-gray precipitate:

$$SnCl_2 + 2\ HgCl \longrightarrow SnCl_4 + 2\ Hg$$

EXERCISE 104

A STUDY OF LEAD AND SOME OF ITS COMPOUNDS

Apparatus. Blowpipe and charcoal; 5 test tubes.

Materials. 2 g. lead (obtain some scrap lead from a plumber); lead acetate (R.S.); 0.5 g. lead nitrate dissolved in 5 cc. water; ammonium sulfide (R.S.); sulfuric acid; potassium chromate (R.S.); hydrochloric acid; piece of mossy zinc; thread or string about 50 cm. long; 10 cc. hydrogen peroxide.

1. (*a*) Note the physical properties of the metal. Heat a small bit on charcoal. (*b*) Is it easily melted? (*c*) Account for the coating formed on the charcoal.

2. To a piece of mossy zinc about as large as the end of your finger tie a piece of thread and suspend the zinc about the middle of a test tube which is nearly filled with a solution of lead acetate. Set the tube aside until the end of the laboratory period, noting, every few minutes, any changes taking place. Leave the tube and contents until the next laboratory period and again note the changes. (*d*) Describe the results and the chemistry involved.

3. Pour 1 cc. of the solution of lead nitrate into a test tube and add one drop of ammonium sulfide. Black lead sulfide is formed. (*e*) Write the equation for the reaction. Now add to the mixture 7 or 8 cc. of hydrogen peroxide solution, shake the mixture, and set it aside for ten minutes. (*f*) Account for the change in color (recall that hydrogen peroxide is a good oxidizing agent).

4. To the remainder of the lead nitrate solution add 2 or 3 drops of potassium chromate solution. Yellow lead chromate ($PbCrO_4$) is formed. This is used as a pigment under the name of *chrome yellow.*

$$Pb(NO_3)_2 + K_2CrO_4 \longrightarrow PbCrO_4 + 2\ KNO_3$$

EXERCISE 105

SIMPLE CELLS FOR PRODUCING ELECTRIC CURRENTS

Apparatus. Beaker (100-cc.); screw clamp.

Materials. 1 strip each of sheet iron and of zinc about 1 cm. × 10 cm. 50 cc. water, to which is added 1 cc. sulfuric acid.

1. Action of metals in a cell (for producing an electric current). A cell in its simplest form consists of two strips of different metals joined together at one end by a wire while the other ends are immersed in a dilute solution of an acid, such as sulfuric. One of the strips will slowly dissolve in the acid while hydrogen escapes from the surface of the other metal. As a result of the action an electric current is generated and flows

through the wire. To determine which of the metals is dissolved, proceed as follows:

Cut a strip of sheet iron about 10 cm. × 1 cm.; also a similar strip of zinc. Polish one end of each strip so as to remove all foreign matter and then clamp the ends firmly together by a screw clamp. Bend the two strips in the form of the letter V and immerse the disconnected ends in a solution of sulfuric acid (made by adding 1 cc. of the acid to 50 cc. of water), as shown in Fig. 89. (*a*) Note from which metal hydrogen escapes (if the metals are impure, hydrogen may escape from both metals, but the amount escaping from one of the metals will always be in excess). Set the beaker and metals aside for one or two hours (or until the next period). (*b*) Note which of the metals has been dissolved. (*c*) Note the relative positions of zinc and iron in the displacement series.

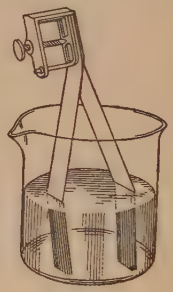

FIG. 89. A simple cell for producing an electric current

When any two metals are connected, as in the above experiment, the one nearest the top of the displacement series is always the one which is dissolved.

2. Galvanized iron consists of strips of iron coated with zinc. If the zinc is worn away at any point so as to expose the iron, and moisture is present, a local current will set up. (*d*) Which of the two metals will be dissolved? (*e*) What is the composition of tin plate? (*f*) If the outside metal is worn away, and moisture is present, which of the two metals will be dissolved?

EXERCISE 106

PAINTS

Apparatus. Blowpipe and charcoal; test tube; funnel and filter paper; small beaker.

Materials. About 5 g. of samples of white paints; 5 cc. carbon tetrachloride; 1 cc. ammonium sulfide (R. S.); filter paper; splint.

1. Introduce 2 or 3 g. of a sample of a white paint into a test tube and then add carbon tetrachloride until the tube is about one fourth full. Shake the tube and contents vigorously; then set it aside for a few minutes (or until you have performed the experiments below). (*a*) Does the solid matter in the paint settle to the bottom of the tube? Shake the tube again, and filter. (*b*) Is the filtrate clear? (*c*) In what form is the solid matter present in the paint? (*d*) What constituent of paint is soluble in carbon tetrachloride? Place 2 or 3 drops of the filtrate on a piece of filter paper and put it aside until the carbon tetrachloride evaporates. (*e*) Does a transparent "spot" remain? (*f*) If so, what is the cause of it?

2. Stir the sample of paint and then place 2 or 3 drops of it on a cavity in the piece of charcoal. Heat this gently in the oxidizing flame of the blowpipe until any oil present burns away; then increase the heat for two or three minutes. Any white lead present is reduced to metallic lead; while any zinc oxide present will form a coating which is yellow while hot, but becomes white on cooling. In this way test samples for white lead and zinc and (*g*) record your results.

3. Coat a wooden splint with white paint and place it in water containing a few drops of ammonium sulfide. (*h*) Interpret your results. (*i*) What would be the result of painting the interior of a chemical laboratory with white-lead paint?

EXERCISE 107

A STUDY OF SOME OF THE COMPOUNDS OF MANGANESE

Apparatus. 6 test tubes.

Materials. 0.1 to 0.2 g. potassium permanganate ($KMnO_4$); crystal of ferrous sulfate; sulfuric acid; ammonium sulfide (R. S.); ammonium carbonate (R. S.); sodium hydroxide; manganese chloride solution (R. S.).

1. (*a*) Examine the physical properties of potassium permanganate. Dissolve about 0.1 g. of it in 5 cc. of water. Add a drop of the solution to a solution of the ferrous sulfate containing 2 or 3 drops of sulfuric acid. The ferrous sulfate is changed to ferric sulfate, the oxygen in the reaction (see equation below) coming from the permanganate, which is a good oxidizing agent.

$$2\ FeSO_4 + H_2SO_4 + O \longrightarrow Fe_2(SO_4)_3 + H_2O$$

2. In potassium permanganate the manganese acts as an acid-forming element. It also acts as a base-forming element in certain compounds. Try the action of ammonium sulfide, ammonium carbonate, and sodium hydroxide, respectively, on a solution of manganese chloride. (*b*) Describe the results and write the equations for the reactions involved.

EXERCISE 108

A STUDY OF SOME OF THE COMPOUNDS OF CHROMIUM

Apparatus. 6 test tubes.

Materials. Solution of potassium chromate (R.S.); lead acetate solution (R.S.); barium chloride solution (R.S.); ammonium sulfide (R.S.); sodium carbonate solution (R.S.); sodium hydroxide; chromium sulfate solution (R.S.).

1. (*a*) *Chromates and dichromates.* Write the formula for potassium chromate; for potassium dichromate. (*b*) Is the chromium an acid-forming or a base-forming element in these

compounds? Add 2 or 3 drops of sulfuric acid to a little potassium chromate solution. (*c*) Explain the results.

2. Try the effect of a solution of potassium chromate on a solution of a compound of lead; also on a compound of barium. (*d*) Describe the results and write the equations for all the reactions involved.

3. *Salts of chromium.* Try the effect of the following reagents on a solution of a salt of chromium: ammonium sulfide, sodium carbonate, sodium hydroxide. (*e*) Describe the results and write the equations for all the reactions.

EXERCISE 109

THE DETECTION OF SILVER, LEAD, AND MERCURY WHEN PRESENT IN THE SAME SOLUTION (Optional)

Apparatus. Two beakers; stirring-rod; funnel; test tubes.

Materials. Solutions of silver nitrate (R.S.), lead acetate (R.S.), and mercurous nitrate (R.S.); hydrochloric acid; filter paper; potassium chromate (R.S.); ammonium hydroxide; nitric acid; hot water; litmus paper (blue).

1. The detection of any one metal becomes more complicated when other metals are present in the same solution. As a rule it is necessary so to treat the mixture as to separate the metals from each other. The principle involved is illustrated in the following experiment.

Pour into a beaker 5 cc. of the silver nitrate solution and 2 cc. each of the solutions of lead acetate and mercurous nitrate. Dilute with water to about 200 cc. Now add hydrochloric acid, drop by drop, with constant stirring. The precipitate formed on stirring settles to the bottom of the beaker. Continue adding the acid until a drop of it no longer causes a precipitate when brought in contact with the clear liquid in the beaker. (*a*) Write the equations for the reactions involved. (*b*) Record the names and formulas of the compounds constituting the precipitate. Filter off the precipitate and fill the filter paper

with boiling water, collecting the filtrate in a beaker. The hot water dissolves the lead chloride so that this compound is present in the filtrate. Test for it by adding to the filtrate a few drops of potassium chromate solution. Yellow lead chromate is precipitated.

It is important that all the lead chloride should be removed from the mixture on the filter paper; hence again fill the filter paper with hot water. Wait until it runs through, and again repeat the process, discarding the filtrates.

(*c*) What is the residue on the filter paper? To this residue add 2 or 3 cc. of ammonium hydroxide and collect the filtrate in a test tube. This filtrate contains the silver chloride dissolved from the residue by the ammonium hydroxide. To prove its presence, neutralize the ammonium hydroxide present by adding nitric acid to the liquid until just acid to litmus paper. The silver chloride is precipitated.

(*d*) What effect did the ammonium hydroxide have upon the color of the residue on the filter paper? This change in color is due to the action of ammonium hydroxide on the mercurous chloride, and serves as a test for the presence of the latter.

(*e*) Supposing that the original solution contained only one or two of the metals of the group, how would the absence of the remaining ones be indicated?

EXERCISE 110

BORAX-BEAD TESTS (Optional)

Apparatus. Platinum wire.

Materials. Borax (R. S.); small piece (size of a pin's head) of a compound of each of the following metals: nickel, iron, manganese, copper.

1. Recall the effect of adding a trace of cobalt nitrate to a borax bead (Exercise 72). Repeat the experiment, substituting for the cobalt nitrate, salts of the following metals: nickel, iron, manganese, copper. (*a*) Record your results.

APPENDIX

THE METRIC SYSTEM

With four or five exceptions this system is now used in all civilized countries. The United States and Great Britain are among the few countries that have not formally adopted it, but even in these countries the system is universally used by scientists and is coming into use more and more by manufacturers.

In the metric system each unit is ten times as large as the next lower unit; hence the system is often termed the "decimal system."

1. Length. The unit is the *meter*. It is equal to 39.37 inches.

10 millimeters (mm.) = 1 centimeter (cm.)
10 centimeters = 1 decimeter (dm.)
10 decimeters = 1 meter (m.)
1000 meters = 1 kilometer (km.)

The only measures of length ordinarily used by the chemist are the millimeter and the centimeter; thus, the height of the barometer at the sea level is recorded as 76 cm. (or more commonly as 760 mm.) and not 7 dm. and 6 cm.

2. Volume. The unit generally used is the *cubic centimeter*.

1000 cubic millimeters = 1 cubic centimeter (cc.)
1000 cubic centimeters = 1 cubic decimeter = 1 liter
1000 cubic decimeters = 1 cubic meter

The chemist uses only the cubic centimeter and the liter as measures of volume. Thus, the volume of a test tube is given as (say) 25 cc.; that of a flask as (say) 500 cc., or $\frac{1}{2}$ liter.

3. Weight. The unit is the *gram*. This is approximately the weight of 1 cc. of pure water at its temperature of greatest density (4°). It is equal to 15.43 grains.

10 milligrams (mg.) = 1 centigram (cg.)
10 centigrams = 1 decigram (dg.)
10 decigrams = 1 gram (g.)
1000 grams = 1 kilogram (kg.)

The gram and kilogram are the units of weight most generally used by the chemist. Thus, the weight of a crucible is given as (say) 10.532 g. and not 10,532 mg. or 10 g. 5 dg. 3 cg. 2 mg.

RELATION BETWEEN ENGLISH AND METRIC CONSTANTS

1 pound (troy) = 373.24 grams
1 ounce (troy) = 31.10348 grams
1 pound (avoirdupois) = 453.59 grams
1 ounce (avoirdupois) = 28.3495 grams
1 kilogram = 2.67923 pounds (troy)
1 kilogram = 2.20462 pounds (avoirdupois)
1 liter = 1.05668 United States quarts
1 gallon = 3.78543 liters
1 cubic centimeter = 0.0610 cubic inch
1 cubic inch = 16.3872 cubic centimeters
1 cubic foot = 28,320 cubic centimeters
1 centimeter = 0.3937 inch
1 meter = 39.37 inches

Also note that

1 centimeter = nearly $\frac{2}{5}$ inch
1 meter = nearly 1.1 yards
1 kilogram = nearly $2\frac{1}{3}$ pounds avoirdupois

TABLE OF SOLUBILITIES OF SOME OF THE COMPOUNDS OF THE METALS

	Acetate	Bromide	Carbonate	Chlorate	Chloride	Chromate	Hydroxide	Iodide	Nitrate	Oxide	Phosphate	Silicate (ortho)	Sulfate	Sulfide
Aluminium	W	W		W	W		A	W	W	A	A	A	W	A
Ammonium	W	W	W	W	W	W	W	W	W		W		W	W
Barium	W	W	A	W	W	A	W	W	W	A	A	A	I	W
Calcium	W	W	A	W	W	X	W	W	W	X	A	A	X	X
Cobalt	W	W	A	W	W	A	A	W	W	A	A	A	W	A
Copper	W	W	A	W	W	W	A	W	W	A	A	A	W	A
Ferric	W	W		W	W	W	A	W	W	A	A	A	W	A
Ferrous	W	W	A	W	W		A	W	W	A	A	A	W	A
Lead	W	X	A	W	X	A	A	X	W	A	A	A	I	A
Magnesium	W	W	A	W	W	W	A	W	W	A	A	A	W	A
Manganese	W	W	A	W	W	W	A	W	W	A	A	A	W	A
Mercuric	W	W	A	W	W	X		A	W	A	A		X	I
Mercurous	W	A		W	A	A		A	W	A	A		X	
Nickel	W	W	A	W	W	A	A	W	W	A	A	A	W	A
Potassium	W	W	W	W	W	W	W	W	W	W	W	W	W	W
Silver	W	I	A	W	I	A		I	W	A	A		X	A
Sodium	W	W	W	W	W	W	W	W	W	W	W	W	W	W
Stannic	W	W			W		A	W		A	A			A
Stannous	W	W			W	A	A	W		A	A		W	A
Zinc	W	W	A	W	W	W	A	W	W	A	A	A	W	A

W, soluble in water.

A, insoluble in water; soluble in either HCl or HNO_3 or in both.

I, insoluble in water and in acids.

X, slightly soluble in water and slightly or readily soluble in acids.

TREATMENT IN CASE OF ACCIDENT

Every laboratory should be supplied with the materials necessary for the treatment of cuts and burns. Such wounds when *immediately and intelligently* treated give little or no trouble, but if left to take care of themselves infection may occur and serious results follow. In case of severe wounds the student, after the preliminary treatment in the laboratory, should be sent to a physician.

The following materials required for the treatment of wounds should be kept in the laboratory in a tight cabinet or cupboard:

1. Plain, sterile gauze bandage: 6 rolls, 1 in. × 10 yd.; 6 rolls, 2 in. × 10 yd.
2. Adhesive plaster for holding gauze in place: 2 rolls, 0.5 in. × 10 yd.
3. 1000 cc. of benzine (low-boiling gasoline).
4. 2000 cc. of Seiler's solution (this may be made from tablets purchased at any drug store).
5. Boric acid solution prepared by dissolving 20 g. of boric acid in 1000 cc. of water.
6. Iodine solution—ordinary tincture of iodine (this may be purchased at any drug store).
7. Agnew's solution prepared after the following prescription:

Tannic acid	2 grams
Borax	2 grams
Glycerin	11 milliliters
Camphorated water	q. s. to make 90 milliliters

Different wounds should be treated as follows:

1. *Cuts from glass.* Remove any dirt or grease from the wound and surrounding skin by washing with gauze saturated with benzine. A rather free bleeding at first will help to prevent infection. Finally, wash the wound with a piece of gauze saturated with the iodine solution; then bandage so as to prevent contamination.
2. *Burns. Immediately* apply gauze saturated with Seiler's solution and then bandage.
3. *Burns from acids.* Treat the same as ordinary burns, as directed under 2.
4. *Burns from alkalies. Immediately* wash with a large volume of water and then with boric acid solution on gauze; then bandage.

5. *Acid in eyes. Immediately* wash the eyes with a large volume of water, then with a solution of boric acid. Finally, drop into the eyes 3 or 4 drops of Agnew's solution and then apply to the eyes a small gauze pad which is kept saturated with ice water.

INFORMATION REGARDING APPARATUS AND CHEMICALS

The lists following include the apparatus and chemicals required for the experiments in this notebook. It is always best to furnish each student with as complete an outfit as possible and to hold him responsible for it. Certain pieces may, however, be used in common by a number of students, and these have been placed in a separate list. It is always cheapest to purchase the apparatus and chemicals in as large quantities as possible. The amounts of most of the chemicals needed for a class of ten are so small that their cost will be proportionately much greater than when larger quantities are ordered. *It is always best to order the definite amounts of chemicals listed in the catalogues, such as 100 g. or 1 lb.; otherwise the cost of weighing out odd quantities and preparing these for shipment may amount to more than the cost of the chemicals.* The supplies may be obtained from any of the large dealers. Catalogues will be sent on application and should be in every school. The following are the addresses of some of the largest firms: Burrell Technical Supply Co., Pittsburgh, Pa.; Cambridge Botanical Supply Company, Waverley, Mass.; Central Scientific Company, Chicago, Ill.; The Chemical Rubber Co., Cleveland, Ohio; Chicago Apparatus Co., Chicago, Ill.; The Fisher Scientific Co., Pittsburgh, Pa.; Kauffman-Lattimer Co., Columbus, Ohio; L. E. Knott Apparatus Co., Boston, Mass.; Eimer and Amend, New York City; E. H. Sargent & Co., Chicago, Ill.; Arthur H. Thomas Company, Philadelphia, Pa.; W. M. Welch Mfg. Co., Chicago, Ill.; The Will Corporation, Rochester, New York.

A list of the supplies needed should be sent to a number of firms for quotations on prices. In ordering any piece of apparatus a certain form in some catalogue should be designated; otherwise it will be impossible to compare the prices. In general it is best to purchase as simple a form of apparatus as possible; for example, 20 cents will buy a Bunsen burner which for ordinary

purposes is preferable to those costing $1. A person experienced in the purchase of supplies will always find it possible to reduce materially the cost of the order.

It is generally cheaper to purchase from local stores ordinary pieces of apparatus and materials commonly sold by such stores, rather than to order them through supply houses. Thus, an iron spoon can be purchased at any local five-and-ten-cent store and will cost less than when purchased through a supply house. Similarly, common salt can be obtained at a lower cost at the corner grocery than through a supply house. A few substances are listed that are not ordinarily furnished by supply houses. Thus, such houses do not ordinarily list "natural rubber," but this can be purchased at a trifling cost from any manufacturer of rubber tires or other rubber commodities. Denatured alcohol may be used where alcohol is specified, but it is better to use the pure alcohol. This can be secured tax free from supply houses by presenting a permit secured from the Federal prohibition officer of your state. *Such alcohol must be kept by the instructor and given out to students only on signed blanks giving the use for which the alcohol is desired.*

Materials purchased of supply houses should be ordered at least three or four months in advance, for the rate of delivery is often very slow. It is best for the teacher, during the spring months, to make up and place the entire order for the following academic year. *The teacher should be careful in giving exact specifications and should insist on getting what he orders. All materials not in accord with the specifications should be promptly returned; otherwise his laboratory may become the dumping ground for poor or off-sized materials. It makes a wide difference in the student's work whether or not he is compelled to use (say) glass tubing that is off size and corks that do not fit. The right size costs no more. Be sure to specify it and then insist on getting it.*

APPARATUS REQUIRED FOR EACH STUDENT (TO BE KEPT IN STUDENT'S LOCKER)

Beakers, nest of seven, from 100-cc. to 700-cc.
Blowpipe, mouth.
Bottles, wide-mouthed: one, 60-cc.; five, 250-cc.

Burner, wing-top, for bending glass tubing (Fig. 12).
Calcium chloride drying-tube, straight, 15 cm. in length (Fig. 37, *B*).
Charcoal, 1 piece about 8 cm. × 3 cm. × 2 cm.
Copper gauze, 10 cm. × 12 cm., from 60 to 70 mesh.
Deflagrating-spoon.
Dish, lead, diameter about 6 cm., depth 3 cm.
Evaporating-dish, diameter 7 cm.
Files: round, about 15 cm. in length; triangular, about 15 cm. in length.
Filters, 25, diameter about 11 cm.
Florence flasks: two, 250-cc.; one, 500-cc.
Forceps (steel), 10 cm. long.
Funnel, diameter about 6.5 cm.
Funnel tube, external diameter of tube 6 mm.
Glass tubing: 300 g., soft, external diameter 6 mm., walls 1 mm. thick; 1 piece, hard-glass, 30 cm. in length, internal diameter 1.5 cm., walls 1.8 mm. thick; 1 piece, hard glass, 25 cm. in length and 6 mm. internal diameter.
Glass rod, diameter 2 mm., 1 piece 30 cm. in length.
Medicine dropper.
Mortar (diameter about 8 cm.) and pestle (both of porcelain).
Pipestem triangle for holding porcelain crucible (Fig. 30).
Platinum wire, small (No. 28), 5 cm. long, for flame tests.
Porcelain crucible and lid, diameter about 3.5 cm.
Rubber tubing: internal diameter 5 mm., 1 piece 50 cm. in length; soft, pure gum, internal diameter 5 mm., 1 piece 30 cm. in length for connections etc.
Screw clamp for rubber tubing.
Splints (ordinary cigar lighters), 125.
Sponge.
Spoon, ordinary iron, bowl 5 or 6 cm. long.
Stoppers: rubber, one-hole, to fit the hard-glass test tube; rubber, two-hole, to fit wide-mouthed 250-cc. bottle; rubber, two-hole, to fit 250-cc. flask; rubber, two-hole, to fit 1000-cc. narrow-mouthed bottle.
Stoppers, rubber, 2 one-hole, to fit the hard-glass tubing specified above.
Test tube, graduated, 30-cc., about 20 cm. long, with 0.5-cc. graduations.
Test tube, hard-glass, 15 cm. in length, diameter about 1.8 cm.
Test tubes, 12, length 15 cm., diameter about 1.5 cm.
Test-tube brush.
Test-tube rack.
Towel.

Tripod.
Watch glass, diameter about 8 cm.
Window glass, 4 pieces 10 cm. square.
Wire gauze, 2 pieces 12 cm. square.

APPARATUS TO BE LEFT ON EACH DESK

Bunsen burner, with 75 cm. of rubber tubing to fit.
Clamp, iron, large, for holding flasks and condensers.
Iron tripod (Fig. 58).
Pneumatic trough (Fig. 28). The trough should be about 12 to 15 cm. deep, and large enough to hold 4 or 5 wide-mouthed bottles (250-cc.). It may be round or rectangular. A pan made of granite ware or an earthen crock serves well, or any tinsmith can readily make suitable troughs of galvanized iron.
Ring stand and 3 rings.
Sand bath, iron, about 12 cm. in diameter.

REAGENTS ON EACH DESK

250-cc. bottles filled with the reagents named below. The bottle containing the sodium hydroxide should have ordinary corks, the others should be glass-stoppered.

Ammonium hydroxide (density 0.90).
Hydrochloric acid (density 1.2).
Nitric acid (density 1.4).
Sodium hydroxide solution (10 g. in 100 cc. of water).
Sulfuric acid (density 1.84).

GENERAL APPARATUS FOR TEN STUDENTS

Apparatus marked with a star (*) are required for optional experiments.

*2 sets apparatus for testing conductivity of solutions (Fig. 62); this may be made or may be purchased of supply houses.
*1 Babcock milk-test apparatus complete, either two- or four-bottle (Fig. 76).
1 balance, weighing from 0.5 g. to 500 g., with accompanying weights.
2 balances, sensitive to 1 cg. and made to carry a load of 100 g. with accompanying weights (each set from 1 cg. to 50 g. in covered wooden box).
1 barometer.
5 one-liter, narrow-necked bottles.
1 bottle or flask, 2000-cc. (Fig. 43, *B*).
4 burettes, 50-cc., graduated in 0.1 cc. (Fig. 61).

5 pieces cobalt glass 10 cm. square (for flame tests).
*4 condensers (Liebig), with rubber tubing.
2 sets cork borers (6 in a set), and sharpener.
2 gross corks, best grade, sizes 7, 8, 9, 10, and 12.
1 cylinder, graduated, 100-cc.
1 cylinder, graduated, 500-cc.
1 distilling apparatus for preparing distilled water.
1 magnifying glass, small.
1 microscope, eyepiece 1 inch, objectives $\frac{2}{3}$ and $\frac{1}{6}$.
5 graduated pipettes (5 cc., graduated in divisions of 0.1 cc. or 0.2 cc.).
*1 spectroscope (Fig. 84).
5 thermometers, graduated from $-10°$ to $150°$ C.
5 retorts, glass-stoppered, 150-cc. (Fig. 65).
5 porcelain boats about 8 cm. in length and 6 or 7 mm. wide (these must be of such a size as to slip easily into the hard-glass tubing, specified under the heading "Apparatus required for Each Student").

CHEMICALS ON REAGENT SHELF (FOR USE OF ALL STUDENTS)

The bottles containing solutions should be glass-stoppered. Gummed letters of the alphabet, of different sizes, may be obtained at a little cost from the Tablet and Ticket Co., Chicago, Ill., and these may be used in making the labels for the bottles; or the complete labels for all the common chemicals may be purchased from supply houses. If the class is small, bottles holding 250 cc. will ordinarily serve; if the class is large, then it is better to use bottles holding at least 500 cc. A few of the reagents, such as limewater, are used so extensively that it is better to use a 1000-cc. bottle. Distilled water must be used in making all solutions. A 10 per cent solution signifies 10 g. dissolved in 100 cc. of water.

Acetic acid (36 per cent).
Aluminium sulfate (10 per cent solution).
Ammonium carbonate (25 g. of the solid dissolved in 70 cc. of water and 10 cc. of ammonium hydroxide (density 0.90), and the solution diluted to 100 cc. with water).
Ammonium chloride (10 per cent solution).
Ammonium molybdate solution (obtained directly from dealers).
Ammonium sulfide solution (obtained directly from dealers).
Barium chloride (10 per cent solution).

Borax (solid).
Calcium chloride (10 per cent solution).
Carbon tetrachloride.
Chloroform.
Chlorine water (water saturated with chlorine).
Chromium sulfate (10 per cent solution).
Cobalt nitrate (5 per cent solution).
Copper sulfate (10 per cent solution).
Disodium phosphate (10 per cent solution).
Fehling's solution: solution A, prepared by dissolving 17.5 g. of copper sulfate crystals in 250 cc. of water; solution B, prepared by dissolving 87.5 g. of sodium potassium tartrate in 250 cc. of 10 per cent sodium hydroxide solution.
Ferric chloride (10 per cent solution).
Ferric sulfate (10 per cent solution).
Hydrogen sulfide solution. Prepared by passing hydrogen sulfide slowly through water until the water is saturated with the gas. Since the solution loses its strength after a few days, it is necessary to prepare it as needed.
Iodine solution. Prepared by dissolving 2 g. of iodine and 10 g. of potassium iodide in 100 cc. of water.
Lead acetate (10 per cent solution). If the solution is not clear, add a few drops of acetic acid and shake the mixture; repeat until the solution becomes clear.
Limewater (saturated solution of calcium hydroxide). Shake 3 or 4 g. of the solid hydroxide with 1 liter of water; then set the mixture aside until the excess of solid settles. Use the clear supernatant liquid.
Litmus solution. Dissolve sufficient litmus (litmus cubes) in the water to impart to it a deep color.
Magnesium sulfate (10 per cent solution).
Manganese chloride (10 per cent solution).
Mercuric chloride (5 per cent solution).
Mercurous nitrate (10 per cent solution).
Phenolphthalein (1 g. dissolved in 200 cc. of alcohol).
Potassium bromide (10 per cent solution).
Potassium chromate (10 per cent solution).
Potassium ferrocyanide (10 per cent solution).
Potassium hydroxide (10 per cent solution).
Potassium iodide (10 per cent solution).
Potassium sulfocyanate (10 per cent solution).
Silver nitrate (4 per cent solution).

Starch solution. Prepared by rubbing to a thin paste 4 or 5 g. of starch with cold water and then adding, 3 or 4 drops at a time and with stirring, to 1 liter of boiling water. Add also about 10 g. of zinc chloride (this acts as a preservative). Mix thoroughly, set the mixture aside, and use the clear supernatant liquid.

Sodium carbonate (10 per cent solution).

Sodium chloride (solid).

Sodium thiosulfate (10 per cent solution).

Zinc acetate (10 per cent solution).

CHEMICALS REQUIRED FOR A CLASS OF TEN

The terms in parentheses after the names of the chemicals refer to the grade of materials to be purchased. The abbreviation c.p. signifies "chemically pure." The weights required in each case are given in the metric system, but there is also added the approximate English equivalent. Those chemicals marked with a star (*) are for optional experiments.

This list does not include substances always easily obtained at home stores, such as baking powders, butter, clay, cloth, fats of various kinds, flour, gasoline, gelatin, ice, iron and copper wire, kerosene, lead, milk, molasses, oleomargarine, salt, soap, starch, sugar, tacks, vinegar, yeast, etc.

	Approximate Amounts
Acid, acetic (36 per cent) (c.p.)	500 g. (1 lb.)
Acid, acetic (glacial)	200 g. (½ lb.)
*Acid, formic (50 per cent)	200 g. (½ lb.)
Acid, hydrochloric (density 1.2) (c.p.)	2 kg. (5 lb.)
Acid, hydrochloric (commercial) for use in preparing carbon dioxide	2 kg. (5 lb.)
Acid, nitric (density 1.4) (c.p.)	3 kg. (7 lb.)
*Acid, oxalic (pure)	200 g. (½ lb.)
Acid, sulfuric (density 1.84) (c.p.)	4 kg. (9 lb.)
*Acid, sulfuric (commercial) (density 1.83) for Babcock milk test	500 g. (1 lb.)
Acid, tannic (commercial)	100 g. (¼ lb.)
Alcohol (95 per cent)	1 liter (1 qt.)
Alum (ammonium) pure	500 g. (1 lb.)
Aluminum (turnings or filings)	30 g. (1 oz.)
Aluminum sulfate (pure, crystals)	500 g. (1 lb.)
Ammonium carbonate (pure)	100 g. (¼ lb.)

	Approximate Amounts
Ammonium chloride (pure)	200 g. ($\frac{1}{2}$ lb.)
Ammonium hydroxide (density 0.90) (c.p.)	2 kg. (5 lb.)
Ammonium molybdate solution	500 g. (1 lb.)
Ammonium nitrate (pure)	100 g. ($\frac{1}{4}$ lb.)
Ammonium sulfate (commercial)	500 g. (1 lb.)
Ammonium sulfide solution	500 g. (1 lb.)
Aniline	100 g. (3 oz.)
Antimony	30 g. (1 oz.)
Arsenic	30 g. (1 oz.)
Arsenic trioxide (arsenious oxide) (commercial)	30 g. (1 oz.)
Barium chloride (c.p.)	100 g. ($\frac{1}{4}$ lb.)
Barium nitrate (crystals)	100 g. ($\frac{1}{4}$ lb.)
Barium sulfate	500 g. (1 lb.)
Benzene	200 g. ($\frac{1}{2}$ lb.)
Bismuth	30 g. (1 oz.)
Bleaching-powder	500 g. (1 lb.)
Boneblack	200 g. ($\frac{1}{2}$ lb.)
Borax (commercial)	500 g. (1 lb.)
Boric acid (pure)	100 g. ($\frac{1}{4}$ lb.)
Cadmium chloride (c.p.)	30 g. (1 oz.)
Calcium carbide	500 g. (1 lb.)
Calcium carbonate (precipitated)	500 g. (1 lb.)
Calcium chloride (fused or granular) for filling drying tubes	1 kg. (2 lb.)
Calcium fluoride (fluorspar)	200 g. ($\frac{1}{2}$ lb.)
Calcium hydroxide (hydrated lime)	500 g. (1 lb.)
Calcium sulfate (plaster of Paris)	1 kg. (2 lb.)
Carbon disulfide (commercial)	100 g. (4 oz.)
Carbon tetrachloride (commercial)	2 kg. (5 lb.)
Charcoal (small pieces)	500 g. (1 lb.)
Chloroform (commercial)	200 g. ($\frac{1}{2}$ lb.)
Chromium sulfate	30 g. (1 oz.)
Cobalt nitrate (pure)	30 g. (1 oz.)
Cobalt chloride	30 g. (1 oz.)
Copper (turnings or scrap)	200 g. ($\frac{1}{2}$ lb.)
Copper foil (thin)	100 g. ($\frac{1}{4}$ lb.)
Copper nitrate (pure)	30 g. (1 oz.)
Copper oxide (black, finely powdered)	100 g. ($\frac{1}{4}$ lb.)
Copper sulfate crystals (c.p.)	200 g. ($\frac{1}{2}$ lb.)
Cottonseed oil	200 g. ($\frac{1}{2}$ lb.)

Cyanamide (commercial)	1 kg. (2 lb.)
Dyes: Gallein, fuchsine, methyl violet, malachite green, Congo red	10 g. of each
alizarin paste (20 per cent)	100 g. ($\frac{1}{4}$ lb.)
Formalin	100 g. ($\frac{1}{4}$ lb.)
Gypsum (crystals)	200 g. ($\frac{1}{2}$ lb.)
Hydrogen peroxide	200 g. ($\frac{1}{2}$ lb.)
Iodine	30 g. (1 oz.)
Iron chloride (ferric) (c.p.)	60 g. (2 oz.)
Iron powder (iron reduced by alcohol)	100 g. ($\frac{1}{4}$ lb.)
Iron sulfate (ferrous)	200 g. ($\frac{1}{2}$ lb.)
Iron sulfide (ordinary lumps for preparing hydrogen sulfide)	1 kg. (2 lb.)
Iron wire (picture-frame wire), No. 0	25 m. (25 yd.)
Junket tablets (obtained from any grocer or druggist)	10 tablets
Lead acetate (sugar of lead) (powdered)	200 g. ($\frac{1}{2}$ lb.)
Lead monoxide (commercial)	100 g. ($\frac{1}{4}$ lb.)
Lead nitrate (pure)	60 g. (2 oz.)
Litmus cubes	60 g. (2 oz.)
Litmus paper (100 strips red, 100 strips blue)	2 tubes of each
Magnesium carbonate (powdered)	100 g. ($\frac{1}{4}$ lb.)
Magnesium sulfate (Epsom salts)	500 g. (1 lb.)
Magnesium wire or ribbon	30 g. (1 oz.)
Manganese chloride	60 g. (2 oz.)
Manganese dioxide (pure, powdered)	1 kg. (2 lb.)
Marble (pieces size of a walnut)	2 kg. (5 lb.)
Mercuric chloride (corrosive sublimate)	30 g. (1 oz.)
Mercuric nitrate (c.p.)	30 g. (1 oz.)
Mercuric oxide	30 g. (1 oz.)
Mercurous nitrate (c.p.)	100 g. ($\frac{1}{4}$ lb.)
Mercury	30 g. (1 oz.)
Nickel nitrate	30 g. (1 oz.)
Paraffin	500 g. (1 lb.)
Phenolphthalein	30 g. (1 oz.)
Phosphorus (white)	20 g. (1 oz.)
Photographic developer	1 tube
Portland cement	1 kg. (2 lb.)
Potassium bitartrate (cream of tartar)	100 g. ($\frac{1}{4}$ lb.)
Potassium bromide (granular, pure)	60 g. (2 oz.)
Potassium carbonate (c.p.)	60 g. (2 oz.)

Potassium chlorate (small crystals)	500 g. (1 lb.)
Potassium chloride (c.p.)	100 g. (¼ lb.)
Potassium chromate or sodium chromate (pure)	100 g. (¼ lb.)
Potassium chromium sulfate (chrome alum)	100 g. (¼ lb.)
Potassium dichromate (pure)	100 g. (¼ lb.)
*Potassium ferrocyanide (c.p.)	100 g. (¼ lb.)
Potassium hydroxide (sticks)	100 g. (¼ lb.)
Potassium iodide (pure)	60 g. (2 oz.)
Potassium nitrate (pure)	200 g. (½ lb.)
Potassium permanganate (pure)	30 g. (1 oz.)
Potassium sulfate (pure, anhydrous)	100 g. (¼ lb.)
*Potassium sulfocyanate (c.p.)	30 g. (1 oz.)
Rubber (natural)	500 g. (1 lb.)
Silver nitrate	30 g. (1 oz.)
Soda lime (granular)	500 g. (1 lb.)
Sodium	30 g. (1 oz.)
*Sodium acetate (fused)	500 g. (1 lb.)
*Sodium benzoate (pure)	30 g. (1 oz.)
Sodium bicarbonate (baking-soda)	500 g. (1 lb.)
Sodium bromide (crystals) (pure)	200 g. (½ lb.)
Sodium carbonate (pure, anhydrous)	200 g. (½ lb.)
Sodium hydrogen phosphate (disodium phosphate) (c.p.)	100 g. (¼ lb.)
Sodium hydroxide (sticks)	1 kg. (2 lb.)
Sodium hydroxide (normal solution)	1 liter
Sodium iodide (crystals) (pure)	100 g. (¼ lb.)
Sodium nitrate (pure)	200 g. (½ lb.)
Sodium potassium tartrate (Rochelle salts) (powdered)	500 g. (1 lb.)
Sodium silicate solution (water glass)	500 g. (1 lb.)
Sodium sulfate (crystals) (Glauber's salt)	200 g. (½ lb.)
Sodium thiosulfate	500 g. (1 lb.)
Sulfur (flowers)	500 g. (1 lb.)
Sulfur (roll)	500 g. (1 lb.)
Tartar emetic (potassium antimonyl tartrate)	60 g. (2 oz.)
Tin (mossy)	60 g. (2 oz.)
Zinc (mossy, arsenic-free)	1 kg. (2 lb.)
Zinc (sheet)	200 g. (½ lb.)
Zinc chloride	100 g. (¼ lb.)
Zinc sulfate (crystals)	100 g. (¼ lb.)
*Zinc sulfide	30 g. (1 oz.)

TABLE OF CONSTANTS

LIST OF THE COMMON ELEMENTS, THEIR SYMBOLS, AND THEIR ATOMIC WEIGHTS

O = 16

Aluminum	Al	26.97	Iodine	I	126.932
Antimony	Sb	121.77	Iron	Fe	55.84
Arsenic	As	74.96	Lead	Pb	207.20
Barium	Ba	137.37	Magnesium	Mg	24.32
Bismuth	Bi	209.00	Manganese	Mn	54.93
Boron	B	10.82	Mercury	Hg	200.61
Bromine	Br	79.916	Nickel	Ni	58.69
Cadmium	Cd	112.41	Nitrogen	N	14.008
Calcium	Ca	40.07	Oxygen	O	16.000
Carbon	C	12.000	Phosphorus	P	31.027
Chlorine	Cl	35.457	Potassium	K	39.096
Chromium	Cr	52.01	Silicon	Si	28.06
Cobalt	Co	58.94	Silver	Ag	107.880
Copper	Cu	63.57	Sodium	Na	22.997
Fluorine	F	19.00	Sulfur	S	32.064
Gold	Au	197.2	Tin	Sn	118.70
Hydrogen	H	1.008	Zinc	Zn	65.38

TENSION OF AQUEOUS VAPOR AT VARIOUS TEMPERATURES, EXPRESSED IN MILLIMETERS OF MERCURY

Temperature	Pressure	Temperature	Pressure
15°	12.78	21°	18.62
16°	13.62	22°	19.79
17°	14.52	23°	21.02
18°	15.46	24°	22.32
19°	16.56	25°	23.69
20°	17.51	100°	760.00

WEIGHT IN GRAMS OF 1 LITER OF VARIOUS GASES MEASURED UNDER STANDARD CONDITIONS

Acetylene	1.1621	Hydrogen sulfide	1.5392
Air	1.2930	Methane	0.7168
Ammonia	0.7708	Nitric oxide	1.3402
Carbon dioxide	1.9768	Nitrogen	1.2506
Carbon monoxide	1.2504	Nitrous oxide	1.9777
Chlorine	3.214	Oxygen	1.4290
Hydrogen	0.08987	Sulfur dioxide	2.9266
Hydrogen chloride	1.6398		

TABLE OF SOLUBILITY OF VARIOUS SOLIDS

Substance	Formula	Weight dissolved by 100 cc. of Water at		
		0°	20°	100°
Calcium chloride	$CaCl_2$	59.5 g.	74.5 g.	159.0 g.
Sodium chloride	NaCl	35.70 g.	36.0 g.	39.80 g.
Potassium nitrate	KNO_3	13.30 g.	31.6 g.	246.0 g.
Copper sulfate	$CuSO_4$	14.30 g.	21.7 g.	75.4 g.
Calcium sulfate	$CaSO_4$	0.759 g.	0.203 g.	0.162 g.
Calcium hydroxide	$Ca(OH)_2$	0.185 g.	0.165 g.	0.077 g.

DISPLACEMENT (ELECTROCHEMICAL) SERIES

1. Potassium
2. Sodium
3. Lithium
4. Calcium
5. Magnesium
6. Aluminium
7. Manganese
8. Zinc
9. Chromium
10. Iron
11. Cobalt
12. Nickel
13. Tin
14. Lead
15. **Hydrogen**
16. Copper
17. Arsenic
18. Bismuth
19. Antimony
20. Mercury
21. Silver
22. Platinum
23. Gold

Thurs Jan 23rd
Thurs Feb
Thurs
Thurs
Thurs
Thurs Mar

Thurs April

Thurs May

Y0-BWB-331